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CHEMISTRY

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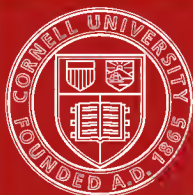
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ELEMENTARY CHEMISTRY

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ELEMENTARY CHEMISTRY

INORGANIC AND ORGANIC

ADAPTED TO THE REQUIREMENTS OF THE 'ALTERNATIVE'
ELEMENTARY SYLLABUS OF THE SCIENCE
AND ART DEPARTMENT.

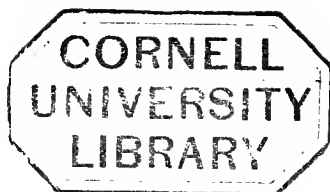
BY

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PREFACE.

THIS LITTLE WORK has been written primarily for young students who are preparing for the examinations in Elementary Chemistry, on the lines of the new 'Alternative' Syllabus of the Science and Art Department ; but, being essentially a brief and simple outline of the Chemistry of Common Things, it is hoped that it may be found useful to all who wish to acquire a knowledge of the elementary principles of Chemistry, so far as concerns the common objects and phenomena of every-day life.

It is also hoped that this little book may serve as an introduction to the study of other branches of natural science, such as Botany, Physiology, Geology, Astronomy, Electricity, &c. It is impossible to properly understand many of the facts and principles of these sciences without at least such a knowledge of Chemistry as this book supplies ; but as an extensive acquaintance with the subject is not necessary for this purpose, the author has endeavoured to avoid the use of technical terms, and to employ simple language throughout.

An endeavour has also been made to render the book thoroughly practical in its character ; and therefore, although the author has written primarily for students who are attending classes under the guidance of a teacher, yet such simple instructions are given, especially in the earlier pages, that even

the self-taught student may find no difficulty in the performance of the various experiments.

The subjects embrace the whole of the Syllabus of the Science and Art Department, and are, with a few exceptions, arranged in the same order. A few additional subjects have also been introduced with the hope of increasing the usefulness of the book. Thus, a lesson on 'Other Useful Metals,' and another on 'Coal, Coal Gas, and Flame,' have been added.

Each lesson concludes with a Summary of its chief teachings ; and it is intended that this shall serve to fix on the memory the main facts of the lesson, and recall the various experiments in illustration thereof.

Several of the cuts have been engraved specially for this work, and the others have been selected from Reynolds' 'Experimental Chemistry,' Jago's 'Inorganic Chemistry,' Miller's 'Elements of Chemistry,' Miller's 'Inorganic Chemistry,' Tilden's 'Practical Chemistry,' and Wright's 'Sound, Light, and Heat.'

W. S. F.

EAST DULWICH : *January* 1888.

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ELEMENTARY CHEMISTRY.

PART I.—INORGANIC CHEMISTRY.

LESSON I.

SOLUTION, CRYSTALLISATION, AND FILTRATION.

1. Solution.

Experiment 1.—Place a piece of loaf sugar on wire gauze suspended by threads, and pass it into a glass of water. The sugar gradually disappears, and is diffused through the water. Its presence may be proved by the sweet taste imparted to all parts of the liquid. It has become liquid in consequence of an affinity between its particles and those of the water, and is said to be **dissolved**. Such a union of sugar and water is called a **solution** of sugar ; the water is termed the **solvent** of the sugar, which is said to be **soluble** in water.

2. There is no loss of weight by solution.

Experiment 2.—Weigh carefully one pound of water, and then one ounce of common salt. Mix the two substances in a glass vessel, and stir till the whole of the salt is dissolved. Now weigh again, and notice that, although the salt has become



FIG. 1.—ILLUSTRATING EXPERIMENT 1.

invisible, owing to the complete separation of its particles, there is no loss in weight, the solution weighing 1 lb. 1 oz.

3. Evaporation.

Experiment 3.—Pour a little of this solution of salt into a clean watch-glass or evaporating dish, and evaporate, i.e. drive off the water in the form of vapour by applying heat. After a time the whole of the water will have disappeared, leaving in the glass a small quantity of dry salt. By evaporating the whole of the pound of water in a larger vessel we may obtain the whole of the one ounce of salt.

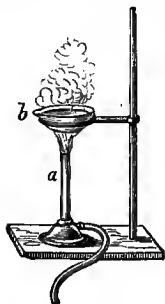


FIG. 2.—APPARATUS FOR EVAPORATION.

a, gas burner (Bunsen's).
b, evaporating dish.

4. Saturation.

Experiment 4.—Add nitre (saltpetre) in the form of a fine powder to cold water, a very little at a time, and keep stirring with a glass rod. The nitre will dissolve in the water. Continue this till it ceases to disappear, and when this is the case the water is said to be **saturated** with nitre, and the mixture is termed a **saturated solution** of nitre in water, for the water has dissolved as much nitre as it can hold in solution, and any more added will remain in the solid state at the bottom of the vessel.

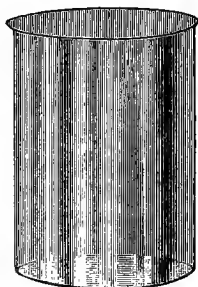


FIG. 3.—A BEAKER GLASS, USED FOR HOT SOLUTIONS.

5. Hot water a better solvent than cold water.

Experiment 5.—Pour this saturated solution of nitre into a thin glass vessel (a beaker), and place the vessel on a support with a spirit lamp or gas flame below it. When hot, add more nitre, and stir. It will be seen that although the solution was saturated when cold, it is now no longer so. Hot water,

therefore, dissolves more nitre than cold water ; and consequently the hot saturated solution will contain more than the cold saturated solution. This is also true concerning most solutions of solids in water.

6. Crystallisation.

There are two slightly different methods of obtaining a salt from its solution in a crystalline form. One is to prepare a solution of the salt, and then evaporate the water by a very gentle heat ; the other is to prepare a *hot* saturated solution of the salt, and simply allow it to cool. In the first case the solution need not be saturated, and it must not be allowed to boil ; for if the temperature is too high, instead of crystals we obtain a shapeless mass. After a certain portion of the water has been driven off by the heat, the solution becomes saturated and then the growth of the crystals commences. The other method of forming crystals from solution is well illustrated by the following experiment :—

7. To form a group of crystals.

Experiment 6.—Make a hot saturated solution of copper sulphate (blue vitriol or bluestone). Suspend an object in it (a piece of coal or a pebble answers very well) by means of a thread, and set it by for some hours. As the liquid cools some of the copper sulphate will separate from the solution in the form of small blue **crystals**, all being of the same shape. These grow as the cooling progresses, for the colder water cannot hold as much in solution as the hot water ; and when thoroughly cooled the sides and bottom of the vessel and the suspended object will be covered with blue crystals, large in proportion to the



FIG. 4.—SHOWING THE MANNER OF PREPARING A GROUP OF CRYSTALS.



FIG. 5.—CRYSTALS OF ALUM.

quantity of the solution. Alum, nitre, or potassium bichromate will also give beautiful groups of crystals if treated in the same manner. When a substance assumes these definite geometric forms it is said to be **crystalline**.

8. Insolubility and suspension.

Experiment 7.—Mix a little finely powdered chalk with water and stir well. The chalk does not dissolve, but remains in the form of small *solid* particles **suspended** in the water. Chalk is therefore said to be **insoluble** in water, and the particles floating about in the liquid are said to be held in **suspension**.

9. Two kinds of impurities in water.

Experiment 8.—Now add to the same mixture of chalk and water a few small crystals of copper sulphate, and stir again. The crystals dissolve, and by the light greenish-blue tint imparted to the water, we may know that the copper sulphate has diffused itself throughout the mass. Thus we produce a sample of water containing two kinds of impurities:—

- (1) **Impurity in suspension**—chalk particles.
- (2) **Impurity in solution**—copper sulphate.

10. Separation of impurities from water.

Experiment 9.—If we set this mixture aside for a little time, we find that the solid chalk particles settle to the bottom of the vessel, while the dissolved copper sulphate remains as a liquid diffused through the water. We may now pour off the clear bluish liquid so gently as to leave the chalk behind, almost free from water and copper sulphate. This is termed **decantation**. By mixing the chalk with a fresh supply of water, and again decanting after the chalk has collected at the bottom, we may get it entirely free from the copper sulphate, and by subsequent drying it may be freed from water. The copper sulphate also may be obtained in a pure state by the evaporation of the water. Thus we are able to separate water from its impurities, and the dissolved from the suspended impurities.

11. Filter and filtration.

Experiment 10.—But there is a better method of separating *suspended* impurities from water than decanting. It is termed **filtration**, and the apparatus employed is called a **filter**. The

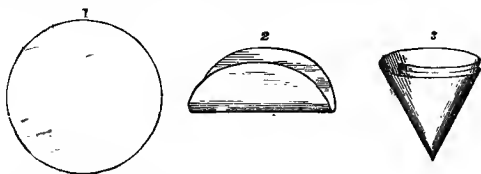


FIG. 6.—SHOWING HOW THE FILTER IS TO BE FOLDED.

simplest filter, and the one generally employed for chemical purposes, consists of a glass funnel (metal funnels are not used, because they are corroded by some of the chemicals), into which is fitted a piece of porous paper. Blotting-paper may be used, but an unglazed paper is manufactured specially for the purpose, generally ready cut in circles. The funnel may be supported by a ring of wire attached to an upright support. The circular paper is first doubled, then doubled again to form a quadrant; it is next opened out to form a conical cup, by taking three thicknesses of paper on one side and one on the other. After placing it in the funnel, a little water will be necessary to make it adhere to the glass. The pores of the paper are so small that only the liquid can pass through, and the solid impurities remain on the filter. The liquid to be filtered should be poured very gently into

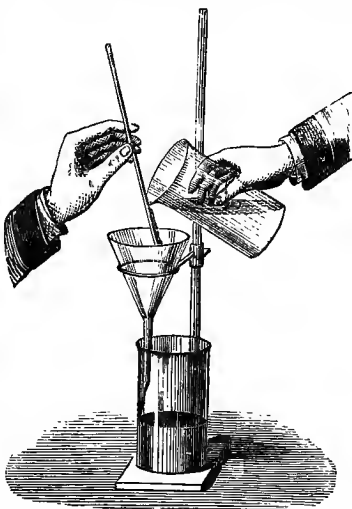


FIG. 7.—APPARATUS FOR FILTRATION.

the filter, so as to avoid breaking the paper. It is a good plan to allow it to run slowly down a glass rod, the lower end of which almost touches the paper. The clear liquid which runs through the filter is termed the **filtrate**.

12. Water will not dissolve all solids.

Experiment 11.—Drop a piece of red sealing-wax into a glass of water and stir. The sealing-wax does not dissolve. In fact, it is absolutely insoluble in water, however long it may be kept in it.

13. Methylated spirit dissolves sealing-wax.

Experiment 12.—Now take the sealing-wax from the water, and put it into a vessel of methylated spirit (methylated alcohol), stirring as before. After a short time the spirit will begin to turn red, and the wax to decrease slightly in size. If allowed to remain in the spirit for several hours it will be first softened and then dissolved. But it will be seen that there is a red sediment at the bottom. The spirit dissolves the gum (shellac) of the sealing-wax, but not the colouring matter (red-lead or vermilion). We thus see that **water is not the only solvent**, and that a substance insoluble in water may be dissolved in some other liquid.

14. Solution and crystallisation of sulphur.

Sulphur (brimstone) is insoluble in water, but it may be rapidly dissolved in a liquid known as carbon bisulphide; and by the evaporation of this solvent (at ordinary temperatures, as it is highly inflammable) the sulphur may be obtained in a crystalline form. The carbon bisulphide has a very unpleasant odour; it will therefore be advisable to allow the evaporation to take place in the open air.

15. Some important solvents.

Here is a list of a few other solids, some insoluble in water, but all soluble in other liquids; and in each case, after dissolving, the dry solid may be again obtained by the evaporation of the solvent:—

Insoluble in water.

Phosphorus,	soluble in	carbon bisulphide.
Gum mastic,	„	alcohol.
Resin,	„	alcohol.
Caoutchouc,	„	carbon bisulphide.

Soluble in water.

Copper chloride,	soluble* in	alcohol.
Calcium chloride,	„	alcohol.

16. Deliquescence.

Calcium chloride is *very* soluble in water. Its affinity for water is so great that, if exposed to the air, it will rapidly absorb moisture from the air sufficient to dissolve itself. If balanced on a pair of scales and then exposed to air, the increase in weight will show the amount of water absorbed. Substances which absorb moisture readily from the air are said to be **deliquescent**. Common salt is slightly, and caustic soda, caustic potash, and calcium chloride, very deliquescent.

17. Salt water is heavier than fresh water.

When water contains a substance in solution, its specific gravity (weight compared with that of an equal volume of pure water) is higher. Thus, sea water is heavier, volume for volume, than fresh water; and the latter may be made to float on the former if poured on very gently.

18. Spirit is soluble in water.

Experiment 13.—Pour water into a tall, narrow glass vessel till half full. Add a little aniline or other soluble colouring matter to some methylated spirit in another vessel. Now pour the coloured spirit *very gently* on the water in the tall vessel. The spirit will float on the water, for it is lighter than water. Now shake the vessel, and it will



FIG. 8.—COLOURED ALCOHOL FLOATING ON WATER.

be seen that the spirit and the water will completely mix, the whole becoming evenly coloured throughout. Hence we say that spirit is soluble in water, or that water is soluble in spirit.

19. Oil is insoluble in water.

Experiment 14.—Pour some olive oil into a vessel of water. The oil, being lighter, floats on the water. Shake the two liquids together vigorously, and the oil may still be seen in the form of minute globules *suspended* in the water. If the mixture be now set at rest for a short time, the oil globules will all rise and collect at the top. Oil and water, therefore, are insoluble in each other.

From the last two and similar experiments we learn that **some liquids are soluble and some are insoluble in one another.**

20. Other illustrations.

The following **experiments** should be attempted by the student, and the results noted :—

- (1) Mix water with sulphuric acid (oil of vitriol).
- (2) Mix mercury (quicksilver) with water.
- (3) Mix mercury with oil.

The first of these illustrates solubility, and the others insolubility.

21. Solubility of gases.

Experiment 15.—Some gases are soluble in certain liquids. Heat some ordinary drinking water or rain water in a flask over a spirit lamp. Long before the water boils, bubbles may be seen rising to the surface. These consist of gases which have been dissolved from the atmosphere by the water, probably for the most part while the water was suspended in the air or falling as rain.

22. Gases soluble in water.

Ammonia gas is exceedingly soluble in water, one volume of water at a low temperature dissolving over 1,000 volumes of the gas. Warmer water cannot contain so much ammonia in

solution ; and, if the solution be boiled, the whole of the gas is driven off. The solution of ammonia is sold by chemists under the name of hartshorn. **Hydrochloric acid** is a gas very soluble in water, and its solution is sold as spirits of salt or muriatic acid. Water will dissolve about its own volume of carbonic acid gas. "**Soda water**" and many other mineral waters contain this gas. Gases held in solution by cold water may always be driven off by heat.

23. Summary.

The following is a summary of the main facts taught in this lesson :—

1. When the particles of a solid substance become thoroughly intermingled with the particles of a liquid, the solid is *dissolved* in the liquid, and the mixture is termed a *solution*.
2. When a substance is dissolved it does not lose in weight.
3. Crystals of a solid may be obtained by evaporating the liquid in which it is dissolved.
4. When a liquid contains as much of a dissolved solid as it can possibly hold in solution, it is said to be *saturated*.
5. Hot water is a better solvent for solids than cold water.
6. Crystals of a solid may be obtained by the cooling of its hot saturated solution.
7. Solid particles floating in a liquid are said to be held in *suspension*, and such particles may be separated by filtration.
8. Dissolved impurities in a liquid may be separated by the evaporation of that liquid.
9. Some solids are insoluble in water, but may be dissolved in other liquids.
10. A liquid has a higher specific gravity when it contains solids in solution than when pure.
11. A deliquescent substance is a solid which readily becomes liquid by the absorption of moisture from the atmosphere.
12. Some liquids are soluble in other liquids.
13. Some gases are soluble in water, but may always be driven off by the application of heat.

These facts should all be committed to memory by the

student, who should also endeavour to recall all the experiments by which they are illustrated, and, if possible, perform them himself.

QUESTIONS ON LESSON I.

1. Explain the terms *solution*, *solvent*, and *soluble*; and illustrate your answer by reference to some simple experiment.
2. How would you obtain crystals of a salt from its solution?
3. What is meant by the term *saturation*? How would you prepare a saturated solution of salt?
4. What is the difference between a hot saturated solution and a cold saturated solution of nitre? How would you obtain crystals of nitre from a solution without resorting to evaporation?
5. Explain the terms *crystal* and *crystallisation*. How would you prepare crystals of alum?
6. What is meant by filtration? Describe a simple form of filter and the method of using it.
7. What do you understand by the terms *suspension* and *solution* as applied to the impurities in water?
8. Chalk and sugar are mixed with water. Describe fully the method by which you would separate the chalk and sugar from each other and from the water.
9. How would you make solutions of sulphur, shellac, phosphorus, resin, and caoutchouc?
10. Name a few substances (solids) which have more than one solvent.
11. Explain the term *deliquescent*. Give examples of deliquescent substances.
12. Describe experiments illustrating the solubility and the insolubility of liquids in each other.
13. Name some gases which are soluble in water. How would you set these gases free from their solutions?

LESSON II.

NATURAL WATERS.

24. Natural waters are never pure.

ABSOLUTELY pure water is never found in nature. The waters of springs, rivers, and seas contain various impurities, dissolved and suspended, derived from the rocks, and from the growth and decay of animal and vegetable substances. The impurities obtained from rocks are termed *inorganic*, and those produced by the growth and decay of animal and vegetable substances, *organic*. Even rain water contains impurities, especially that

falling in large towns, and if evaporated to dryness will sometimes leave a residue of solid matter. We may now give a little attention to the nature of the impurities to be met with in the various natural waters.

25. Rain water.

Rain is formed by the condensation of the pure vapour produced by the evaporation of the water of the ocean, lakes, rivers, &c. Consequently, it will contain no impurities except those derived from the atmosphere; and these are chiefly gases held in solution.

Experiment 16.—Collect some rain water in a clean vessel, transfer it to a glass flask, and apply heat. Bubbles of gas will soon begin to ascend. This gas consists chiefly of—

- (1) Carbonic acid gas, composed of carbon and oxygen.
 - (2) Oxygen
 - (3) Nitrogen
- } elementary or simple gases.

26. Spring and river waters.

These contain, in addition to gaseous impurities, both dissolved and suspended impurities obtained from the rocks over or through which they flow, and varying in nature according to the character of these rocks. Thus, rivers flowing over limestone, or springs fed by water which percolates through limestone, contain more or less lime in solution. Suspended impurities may also be found, some derived from the wearing down of insoluble rock substances, and others of organic origin. The following is an analysis of Thames water:—

Composition of Thames water at London Bridge.

Water	69971.95 grains.
Carbonate of lime	8.12
Chloride of calcium	6.97
Sulphate of soda	3.11
Chloride of sodium	2.38
Sulphate of potash27
Silica12
Chloride of magnesium08
Organic matter in solution	2.34
Organic matter in suspension	4.66
	<hr/> 70000.00 grains (one gallon).

27. Ponds and lakes.

The water of ponds and lakes contains impurities similar to those found in rivers ; but where the pond or lake has no outlet, and great loss of water is produced by evaporation, the proportion of foreign matter may be continually increasing. The following list of natural waters shows how they vary with regard to the amount of matter *in solution* :—

		Grains per gallon.
1. Loch Katrine	water contains	4·2
2. River Jordan	„	7·0
3. River Seine (at Paris)	„	21·0
4. Dead Sea	„	4,000 to 16,000·0

Loch Katrine water rests on a bed of granite, one of the hardest of rocks, and one not easily decomposed by water or atmospheric gases. The Seine and the Thames both flow over softer rocks, which are also more readily decomposed by the gases dissolved in water. The river Jordan is one of the feeders of the Dead Sea. This lake, being much below the level of the neighbouring seas, has no outlet, and the amount of evaporation is great ; hence the large proportion of dissolved matter.

28. Sea water.

The sea is the common reservoir for the impurities of numberless rivers. It is continually receiving the water discharged by these rivers, and is at the same time giving off as much in the form of pure water-vapour by evaporation. We can thus understand how it is that sea water contains as much as 2,400 grains of dissolved matter per gallon, while the rivers flowing into it contain only from 4 to 20 or 30 grains per gallon. The chief substance found dissolved in sea water is common salt (chloride of sodium) ; the others are shown in the following analysis :—

Composition of the water of the English Channel.

Water	67606·6 grains.
Chloride of sodium	1964·2
Chloride of magnesium	256·7
Sulphate of lime	98·5
Chloride of potassium	53·6
Sulphate of magnesia (Epsom salts)	16·1
Carbonate of lime	2·3
Bromide of magnesium	2·0
<hr/>	
70000·0 grains (one gallon).	

29. How to estimate the amount of solid matter in impure water.

If it is required to determine the amount of solid matter in any given quantity of spring, river, or other impure water, evaporate the liquid to dryness, and weigh the residue. If, however, it is required to determine the proportion of suspended and dissolved impurities separately, first filter the water to remove the suspended matter, and then evaporate the liquid which passes through. The suspended and dissolved impurities may now be weighed separately.

30. Hard water.

When water contains a large quantity of lime in solution, it will not form a lather readily with soap, and is then said to be *hard*. **Hardness** is generally due to the presence of bicarbonate of lime. Chalk, limestone, and marble are varieties of carbonate of lime. These rocks are insoluble in *pure* water, but natural waters always contain carbonic acid gas in solution, and water containing this gas will convert the insoluble carbonate of lime into the soluble bicarbonate of lime, thus rendering it hard. Consequently, we can easily soften hard water by ridding it of its carbonic acid gas. This may be done by boiling. Let the hard water be boiled in a clean glass vessel, and as the carbonic acid gas passes off, a deposition of carbonate of lime will take place on the sides and bottom of the vessel. It is the repeated removal of lime in this way from drinking water that produces the 'fur' on our tea-kettles and boilers. Hard water may also be softened by the addition of lime (quicklime). The lime added combines with the carbonic acid gas present, forming carbonate of lime; and this, being insoluble in water, forms a white sediment. The water, deprived of this gas, can now no longer hold in solution the carbonate of lime originally contained in it, and this also is precipitated, leaving the water comparatively soft. Water containing sulphate of lime (gypsum) is said to be **permanently hard**, for it cannot be softened conveniently without the introduction of still more objectionable substances. The waters of both the Thames and the Trent are hard: the hardness in the former is due to carbonate of lime, and it may

therefore be softened ; but in the latter case the hardness is permanent, on account of the presence of gypsum.

31. Mineral waters.

Some spring waters contain so much dissolved mineral matter that they are termed **mineral waters**. The kind and amount of mineral substance present depend, of course, on the nature of the rocks through which the water passes. For instance, the **Epsom springs** contain large quantities of sulphate of magnesia, which has consequently been termed **Epsom salts**. Some mineral waters have a high temperature, and as hot water is a better solvent than cold water, such waters contain much dissolved mineral matter. The boiling springs of Iceland contain a great deal of dissolved silica, a substance which is but sparingly soluble in cold water. When spring water contains very much lime it is said to be *calcareous* ; when rich in silica, *silicious* ; if it contains large quantities of iron it is termed *chalybeate* or *ferruginous*.

32. Summary.

1. Natural waters always contain impurities, and these impurities may be either *organic* or *inorganic*.

2. Rain water contains gaseous impurities derived from the atmosphere.

3. The waters of springs, rivers, and lakes contain, in addition to these gaseous impurities, mineral matter, derived from the rocks over or through which the waters flow, and also organic impurities.

4. Sea water contains about $3\frac{1}{2}$ per cent. of salts in solution, the chief being common salt (sodium chloride).

5. Water containing the salts of lime in solution will not readily form a lather with soap, and is said to be *hard*.

6. Hard water may be softened by boiling, or by the addition of lime.

7. Water is '*permanently hard*' when it contains sulphate of lime.

8. *Mineral waters* contain very large quantities of mineral salts in solution, especially when, as is often the case, they are hot,

QUESTIONS ON LESSON II.

1. What are the chief impurities found in rain water? Whence does it derive these impurities, and how would you remove them?
2. How do spring and river waters differ from rain water?
3. Name some of the substances often present in spring water.
4. How would you account for the small quantity of dissolved matter in the water of Loch Katrine, and for the very large proportion in the water of the Dead Sea?
5. What are the substances found dissolved in sea water, and which of these are most abundant?
6. How can you tell whether water is 'hard' or 'soft'? What is the cause of hardness in water?
7. Describe two methods by which hard water may be rendered soft.
8. When is water said to be permanently hard, and why?
9. How is 'fur' formed in kettles and boilers?
10. What are 'mineral waters'? Give some examples.

LESSON III.

AIR.

33. What is the atmosphere?

The atmosphere is an envelope of gas which completely surrounds the globe. Although invisible, we have many proofs of its existence. For instance, while breathing we can feel its motion through our mouths or our nostrils; and the expansion of our lungs at each inspiration proves that something which **occupies space** has entered. If we take a large sheet of cardboard and wave it in the air, we can feel that something **offers resistance** to its motion. A vessel is said to be empty when it contains nothing but air, but we may easily prove that the vessel is quite full by the following simple experiment.

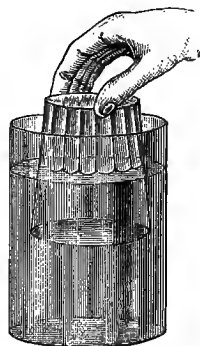


FIG. 9.—EXPERIMENT PROVING THAT AIR OCCUPIES SPACE.

Experiment 17.—Invert a glass tumbler, or, better still, a very deep glass jar, and press it under the surface of water con-

tained in a larger vessel. We now see that the water does not rise and fill the tumbler, which must therefore be full of something occupying space and offering resistance to the water. This is air. Wind is simply air in motion.

34. The volume of air varies with the temperature.

Experiment 18.—Fill a glass vessel with water, invert it, and stand it in a vessel of water without allowing any to escape. Fit a flask with a sound cork through which a bent tube passes into the flask, the other end dipping under the water into the neck

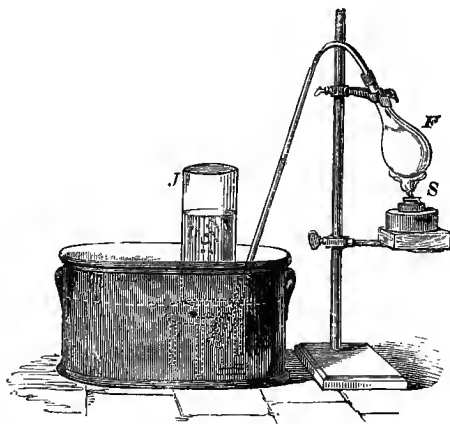


FIG. 10.—APPARATUS FOR ILLUSTRATING THE EXPANSION OF AIR.
F, flask containing air; S, spirit lamp; J, jar containing the air driven out by expansion.

of the first vessel. (See fig. 10.) Now apply heat to the ‘empty’ flask, and bubbles of air will immediately rise and displace the water in the other vessel. The rise in temperature has caused the air to **expand**, and the amount of expansion is rendered visible by the displacement of the water. One-half of the air contained in a glass flask may readily be driven out in this manner. If now we remove the source of heat, the air will quickly cool and **contract**; and, as the air which has been driven out cannot pass back again, water will rush into the flask to take its place, probably causing a fracture, especially if the glass has been

made very hot. If, instead of heating the air by means of a lamp or gas flame, we simply hold the flask in the hand, we shall see that the warmth of the hand is sufficient to cause a certain amount of expansion. We thus learn not only that air expands when its temperature is raised, but that it **expands very readily**, even with only a slight increase in temperature; also that it **contracts readily** when its temperature is lowered.

35. The volume of air is reduced by pressure.

Experiment 19.—Repeat the experiment mentioned in Art. 33, and notice that the water does rise a *little* way inside the tumbler, although none of the air has been allowed to escape. The air has been compressed into a smaller volume by the pressure of the water above the level of the mouth of the tumbler; but as the pressure of this water is only slight, we may say that **air is easily compressible**. The following experiment, however, is much more striking in its result :—

Experiment 20.—Procure a glass tube bent like the letter U, but with one arm much longer than the other (36 inches and 6 inches respectively will answer well), and the short arm closed. (See fig. 11.) Pour mercury into the tube while erect. This will fill the bend at the bottom and prevent any escape of air from the short arm. Continue pouring in mercury, and notice that the higher the column of liquid the smaller the volume of the air in the short arm, for it is compressed by the weight of the mercury, and the pressure exerted by the mercury is proportional to its height above the level of the liquid in the short arm. If the top of the mercury in the long arm is 30 inches higher than the top of the liquid in the short arm, then the air will occupy only about half its original volume. The less the volume of a given quantity of air the greater its **density**. A

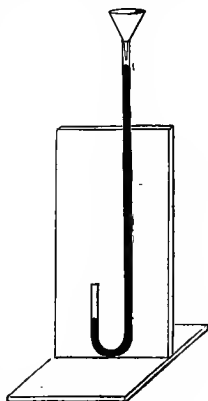


FIG. 11.—ILLUSTRATING THE COMPRESSIBILITY OF AIR.

substance is said to be dense when its particles are closely packed together. The **air on mountains** is less dense than the air at the sea level. This is due to the greater pressure exerted on the latter by the overlying atmosphere. The density of the air decreases so gradually with the altitude that it is difficult to tell accurately the **height of the atmosphere**, but it is generally given as 40 or 50 miles.

36. Air is elastic.

A substance is said to be elastic when it resumes its original form or volume after compression, stretching, twisting, or bending. If we compress air its volume is reduced, but as soon as the pressure is removed it at once resumes its former volume.

Experiment 21.—Close the nozzle of an ordinary syringe, and then press on the air contained in it. If the piston fits well



FIG. 12.—APPARATUS FOR ILLUSTRATING THE COMPRESSIBILITY AND ELASTICITY OF AIR.

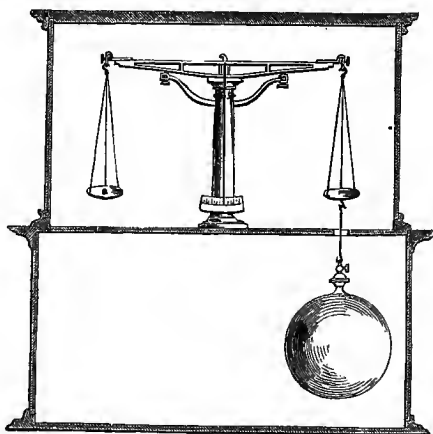


FIG. 13.—APPARATUS FOR WEIGHING AIR AND OTHER GASES.

the air is compressed, and on taking off the pressure the piston will fly back again. Now pull out the piston so as to compel

the air to occupy a larger space (for if the capacity of a vessel is enlarged the air will always diffuse itself equally throughout), and on letting go it will return to its former position.

37. Air has weight.

Experiment 22.—That air has weight may be proved by the following experiment.—Procure an air-tight vessel provided with a stop-cock, weigh it carefully and note the exact weight. Now exhaust the air by means of an air-pump. Weigh again and note the decrease in the weight. This loss of weight will correspond with the weight of the air which the vessel will hold. The weight of 100 cubic inches of air is about 31 grains; 13 cubic feet of air weigh about 1 lb. Air is 800 times lighter than its own volume of water, and 11,000 times lighter than the same volume of mercury. Air, having weight, will exert pressure, and therefore every substance at the surface of the earth has to bear atmospheric pressure.

Experiment 23.—Fill a glass vessel by immersing it under water, and invert it, still keeping its mouth under the surface. Now lift it gradually, till it is almost entirely above the level of the water, and it will remain quite full, for the water is supported by the pressure of the air. The pressure of the atmosphere may thus be made to support a column of water about 34 feet high. Take a glass tube, about 32 inches long, and closed at one end. Fill it with dry mercury,

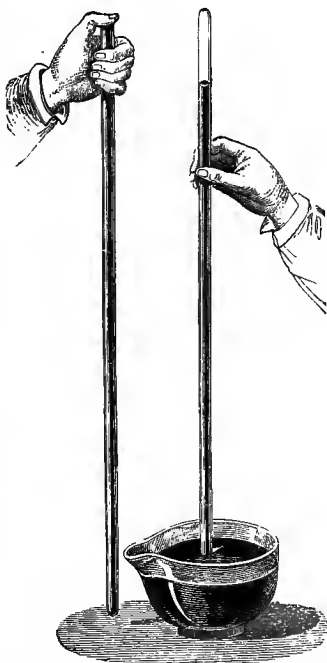


FIG. 14.—ILLUSTRATING THE PRESSURE OF THE ATMOSPHERE.

close it with the thumb, and then invert it in a cup of mercury. On removing the thumb, the mercury will fall till it stands at about $29\frac{1}{2}$ inches above the level of the liquid in the cup. It is the pressure of the atmosphere that supports the column of mercury remaining in the tube.

38. Summary.

In this lesson we learn that air possesses the following properties :—

1. It occupies space.
2. It offers resistance.
3. It easily expands when the temperature is increased, and readily contracts when the temperature is lowered.
4. It is easily compressible.
5. It is elastic.
6. It has weight, and therefore exerts pressure.

Note.—These properties belong not only to the air ; they are **common to all gases**, which may be proved by experiments similar to those given in the foregoing lesson.

QUESTIONS ON LESSON III.

1. How would you prove the existence of an atmosphere surrounding the globe?
2. Describe a simple experiment to prove that air occupies space.
3. What is wind?
4. What is the effect of temperature on the volume of air? Illustrate by an experiment both the expansion and contraction of air.
5. What is the influence of pressure on the volume of a quantity of air? Describe an experiment by which you would reduce its volume one-half.
6. How does the air on mountains differ from that on low grounds? What is the cause of this difference? What is the supposed height of the atmosphere?
7. What is meant by elasticity? How would you prove that air is elastic?
8. How would you prove that air has weight? What is the weight of the air contained in a box 3 feet long, 3 feet wide, and 3 feet high?
9. Describe a simple experiment which proves that the atmosphere exerts pressure.
10. What is the greatest height of a column of water which may be supported by the atmosphere; also of a column of mercury?
11. What are the properties common to gases?

LESSON IV.

COMPOSITION OF THE AIR.

39. Air is composed of two gases.

Experiment 24.—Take a piece of phosphorus¹ about as large as a pea, dry it between blotting paper, and place it on a large piece of dry cork, or in a small porcelain crucible floating in a shallow glass vessel containing a little water. Cover the cork and phosphorus with a stoppered, bell-shaped gas jar (an ordinary bottle will do if the bottom be cut off, and a tight-fitting cork be fixed in the neck). Remove the stopper, and ignite the phosphorus by touching it with a hot wire. Replace the stopper immediately, and watch carefully for any alteration in the level of the water. At first the water in the jar will be driven down by the expansion of the air caused by the heat of the flame, but it will soon begin to rise again, and appear above the level of the water outside.² The jar is now full of white fumes of a new substance, and when the phosphorus has ceased to burn and the jar has become cool, it will be seen that the water inside the jar fills one-fifth of the space previously occupied by air. From this we learn that a certain gas, which supports the combustion of phosphorus, has been taken from the air, and that this gas occupied one-fifth of the volume of the air.



FIG. 15.—AN EXPERIMENT FOR DETERMINING THE COMPOSITION OF THE AIR.

40. Experiment 25.—Allow the apparatus to stand till all the white fumes have disappeared, and we shall find that the

¹ Phosphorus should always be kept under water, and never be touched by the hand, as it ignites at a low temperature. It should be cut while in water, and lifted on the point of a knife.

² If the water around the jar sinks very low, more should be added so as to prevent the entry of air.

jar contains an invisible gas, occupying about four-fifths of the volume of the air with which we started our experiment. Now pour water into the vessel till it reaches the level of that inside the jar, remove the stopper, and introduce a piece of burning phosphorus in a small metal cup. It is immediately extinguished. Thrust a burning match or taper into the jar ; at once it goes out. Here is a gas, then, which is very different from air, for it will not support combustion.

41. Hence we learn that air is composed of two gases : one will support combustion, and forms about one-fifth of the air ; the other will not support combustion, and forms about four-fifths of the air. The former is called **oxygen**, and the latter **nitrogen**. The following table may assist in committing these facts to memory :—

Air		Invisible gas.
		Forms one-fifth of air.
		Supports combustion. 2.
		Invisible gas.
		Forms four-fifths of air.
		Does not support combustion.

42. Use of nitrogen in the atmosphere.

Nitrogen and **oxygen** are both elementary substances.¹ **Nitrogen** is a very inactive gas : it does not combine directly with other elements. Its use in the atmosphere is to modify the action of the oxygen, which, if present alone, would be too powerful for the support of animal life. The particular study of oxygen, however, must be left for our next lesson.

43. An illustration of chemical action.

Experiment 26.—Burn another piece of phosphorus, but this time under a *dry* gas jar, standing on a dry surface. The white fumes are formed as before, but we notice that they gradually collect on the sides of the jar and on the surface at

¹ An element is a simple substance, containing only one kind of matter.

the bottom of it. If after a time we remove the jar, we find that the white substance is solid. By shaking up a little water in the jar we can dissolve this solid; and if we add the solution to a blue vegetable colouring matter, such as a solution of litmus, we see that the colour is changed to red, proving that the solution of the white solid has **acid**¹ properties. Here we have an example of **chemical action**. The phosphorus, which is an elementary substance, combines during its combustion with oxygen of the air, and forms with it a **new and entirely different substance**, called oxide of phosphorus. The two elements are said to have a **chemical affinity** for each other.

44. It is not necessary to ignite the phosphorus in order to make it combine with oxygen; for just as an iron nail will slowly combine with oxygen and form rust or oxide of iron, so phosphorus, exposed to air at ordinary temperatures, will oxidise slowly. Therefore, in analysing² the air, it is not necessary to burn the phosphorus in it, but the phosphorus may be simply suspended in the jar, in which case a very long time must be allowed for the complete removal of the oxygen.

45. The air is a simple mixture.

When two substances combine chemically, a **new substance is produced**, having entirely different properties. This was seen to be the case when phosphorus combined with oxygen. They also combine in certain definite proportions. For instance, a given weight of phosphorus will always combine with exactly the same corresponding weight of oxygen gas. Now, if we mix nitrogen and oxygen gases in the same proportions in which they exist in the air, we produce a mixture exactly corresponding with air in all its properties—a mixture which will support life and combustion just as it is maintained in the atmosphere. But we do not produce a new compound with entirely different properties, for the properties of air are intermediate with those of oxygen and nitrogen—the presence of

¹ See Art. 148 for definition of an acid.

² The splitting of a substance into its elementary constituents is called *analysis*.

nitrogen does not *destroy* the original properties of the oxygen, but simply modifies its action. We therefore regard the air as a simple mixture of oxygen and nitrogen, and not as a chemical combination of these two gases. Again, the **composition of the atmosphere varies slightly** at different times and places, but the composition of a chemical compound never does so.

46. Diffusion of gases.

Nitrogen gas is a little lighter than oxygen. The density of a gas is its weight as compared with that of an equal volume of hydrogen—the lightest gas known. Taking this as 1, the density of nitrogen is 14, and that of oxygen 16. The **density of air** (which, as we have already seen, is composed of about four times as much nitrogen as oxygen) would therefore be—

$$\frac{(4 \times 14) + 16}{5} = \text{about } 14\frac{1}{5}.$$

But if the air is a mixture of two gases of different densities, we should expect to find the heavier gas near the surface of the earth, and the lighter gas floating over it, just as oil will float on water. Is this really the case? In order to answer this question we must resort to experiments. Examine a measured volume of air at or near the sea level, as described in Exp. 24, to ascertain the amount of nitrogen it contains, both air and nitrogen being carefully measured in a graduated vessel. Now repeat the experiment with the air on a high mountain, and we find that, although the air on the mountain is much rarer than that at the sea level, yet the proportion of oxygen to nitrogen in each case is very nearly, if not exactly, the same. This is due to a property called **diffusibility**—a property common to all gases. When we mix oxygen and nitrogen, the two gases will at first have a tendency to arrange themselves one above the other according to their densities; but if we set the mixture aside for some time, we shall find on examination that they have both become completely and evenly diffused throughout the vessel containing them. The winds of course would aid in the admixture of the gases of the atmosphere if they did not possess this property of diffusion, but then in “still” air there would always be a tendency to separation again.

47. Air is necessary for combustion.

Air is necessary not only for the combustion of phosphorus, but of almost all combustible substances.

Experiment 27.—Stand a lighted candle in a shallow vessel, cover with a stoppered bell-jar, and surround with sufficient water to render it air-tight. The candle will burn at first with the same brilliancy and steadiness with which it burns in unconfined air. But as soon as the oxygen has been consumed, its flame is extinguished. If we now remove the stopper of the jar, and pass a bent glass tube under its lower edge so as to allow a free passage of air, the candle will burn as well as in the open air. From this we learn why a lamp or lantern is always provided with openings for the entry of fresh air below the flame, and with one or more openings at the top to allow the heated and spent gas to pass out. A stove, too, must always be so arranged as to allow a free upward passage of air through the fire.

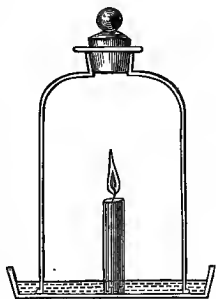


FIG. 16.— EXPERIMENT PROVING THAT AIR IS NECESSARY FOR COMBUSTION.

48. Other gases in the air.

In addition to oxygen and nitrogen, the air contains other gases, the chief of which are **water-vapour** and **carbonic acid gas**. The latter is a heavy gas which will support neither animal life nor combustion. The quantity of it in the air is small, but it is nevertheless an important gas, and its study will form the subject of a future lesson. The water-vapour exists as an invisible gas; its amount is very variable.

49.—The following table gives with more accuracy the percentage composition of the air:—

Composition of the air.

Nitrogen	77·95
Oxygen	20·61
Water-vapour	(average)	.	1·40
Carbonic acid gas	·04
Ammonia gas	.						
Carburetted hydrogen							
Sulphuretted hydrogen							
Sulphurous acid gas							
Nitric acid gas	.						

} Traces only.

100·00

50. Water-vapour in the air.

When a glass of ice-cold water is brought into a warm room, the outer surface of the glass is almost immediately covered with a deposit of little globules of water; and these may soon increase in size till they run in large drops. This water has come from the air, where it existed as an invisible gas or vapour, but the low temperature of the glass has converted it into the liquid state. The drops of dew on vegetation are formed in exactly the same manner when the plants are colder than the surrounding air.

51.—Chloride of calcium and sulphuric acid—the former a white solid and the other a liquid—both absorb water freely when exposed to the air.

Experiment 28.—Place a little of either in a watch-glass or other shallow vessel, and balance in a pair of scales. The gradual falling of the pan containing the substance will prove a corresponding increase in weight, which is due to the absorption of moisture from the air.

52. Evaporation of water.

When an open vessel of water is exposed to the air, the liquid gradually disappears. It has passed off in the form of invisible vapour, which mixes with the air. The water-vapour in the atmosphere is derived from the oceans, &c., in a similar manner by evaporation.

53. Summary.

In the present lesson we have learnt that—

1. The air is composed of two gases, one of which will support combustion, and the other will not.

2. These gases are oxygen and nitrogen, the former constituting about one-fifth of the air, and the latter four-fifths.

3. Nitrogen is an inactive gas, the use of which is to modify the extreme activity of oxygen.

4. When a substance burns it combines with oxygen, and forms a new compound. This combining is termed chemical action.

5. The two gases of the atmosphere form a mixture, and not a chemical combination.

6. All gases are diffusible.

7. Air is necessary for combustion.

8. The air contains water-vapour and small proportions of various other gases.

QUESTIONS ON LESSON IV.

1. Describe fully an experiment which proves that air is composed of two gases.

2. What are the two gases of the atmosphere? Describe the properties of each.

3. What is chemical action? Illustrate your answer by an experiment.

4. For what reasons do we say that air is a *simple mixture* of oxygen and nitrogen?

5. Explain the term diffusibility as applied to gases; and do so with special reference to the gases of the air.

6. What is the density of the air, and what is the density of each of the gases composing it? Explain the term density.

7. Describe an experiment which proves that air is necessary for combustion.

8. What are the gases contained in small quantities in the atmosphere?

9. How would you prove the presence of water-vapour in air?

LESSON V.

OXYGEN AND COMBUSTION.

54. What we have already learnt about oxygen.

We have already ascertained the following facts about oxygen :—

1. It is an invisible gas, forming about one-fifth the volume of the atmosphere.

2. It is a very active gas—a supporter of combustion.

3. When a substance is burnt in air, it combines with oxygen, and as soon as the oxygen is consumed, all burning ceases.

In the present lesson we shall learn how to prepare oxygen, and also study the properties of the gas in a pure state.

55. To prepare oxygen from oxide of mercury.

One of the simplest methods of preparing oxygen in very small quantities is by heating in a test-tube the oxide of mercury (red oxide of mercury or red precipitate).

Experiment 29.—Put a little of this substance into a test-tube of hard glass, and heat it in the flame of a spirit lamp or

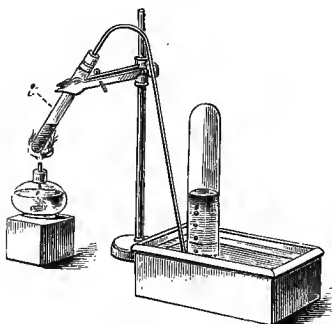


FIG. 17.—APPARATUS FOR PREPARING AND COLLECTING SMALL QUANTITIES OF OXYGEN GAS.

a. Test-tube, containing red oxide of mercury.

Bunsen burner. The red crystalline powder soon turns black, and shortly afterwards a deposit of little globules of the liquid metal mercury may be seen on the cooler part of the tube. If now we thrust into the tube a smouldering taper, or a chip of wood with only a feeble spark at the end, it will immediately burst into a flame, and burn, for a moment, more brightly than it would in air. This proves the presence of a gas in the tube which is a more powerful

supporter of combustion than air itself. This gas is oxygen. It may be collected over water if the test-tube be fitted with an air-tight cork and a suitable delivery tube.

56. Chemical analysis.

This experiment is a good illustration of what is termed chemical analysis. The oxide of mercury is a compound substance containing two elements—mercury and oxygen. When heat is applied, the chemical affinity binding these two elementary substances together is neutralised, and both ele-

ments are set free. The mercury passes off in the form of vapour, which condenses on a cooler part of the tube, and the oxygen becomes an invisible gas. The compound has been analysed or decomposed into its component elements.

57. A more convenient method of preparing oxygen.

Experiment 30.—Oxygen is best prepared for experimental purposes by heating in a glass flask or retort a mixture of chlorate of potash (potassium chlorate) and black oxide of manganese (manganese dioxide). Both these substances contain oxygen, and the gas may be prepared by heating either alone ; but if we mix the two substances, using about equal

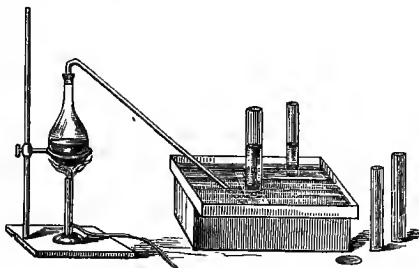


FIG. 18.—APPARATUS FOR PREPARING AND COLLECTING OXYGEN IN LARGER QUANTITIES.

quantities by weight of each, the oxygen comes off at a much lower temperature. The flask must be provided with a cork through which passes a bent glass tube, as shown in fig. 18; or a glass retort may be used instead of the flask. Before applying heat a large vessel of water must be provided, containing one or more jars filled with water and inverted. It is more convenient to have a large metal trough provided with a perforated shelf, on which two or more jars may stand ready for use. Such a vessel is called a **pneumatic trough**. Now allow the delivery tube to dip into the water and apply heat. Almost immediately bubbles of gas will rise through the water ; but these are air bubbles driven out by expansion. Allow this air to escape into the atmosphere and even after oxygen gas begins to come off, it will be better to wait till all the air is driven out. Now allow the gas to pass into each jar until all the water is displaced, then close with a cork or a greased glass disc, and remove from the trough.

58. The combustion of carbon in oxygen.

Experiment 31.—Remove the cork from one of the jars, and allow the gas to remain exposed to the atmosphere ; for oxygen, being heavier than air, will not rapidly diffuse into it. Fasten a few inches of copper wire to the cork, and twist the free end of the wire round a piece of charcoal. Ignite the charcoal and pass it into the oxygen, allowing the cork to *rest lightly* on the mouth of the jar, so as to permit of the escape of the expanded gas. The charcoal immediately begins to glow very brightly, throwing off brilliant sparks, and this continues till most of the oxygen has been consumed. Thus we see that charcoal, which is a variety of the element carbon, burns much more brilliantly in pure oxygen than in air. But we must learn the nature of its combustion by another experiment.

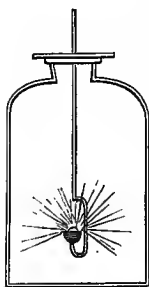


FIG. 19—CHARCOAL BURNING IN OXYGEN.

The Charcoal is supported in a 'deflagrating spoon.'

59. Carbonic acid gas is produced when carbon is burnt in oxygen.

Experiment 32.—Remove the charcoal as soon as the combustion is over, pour in a little clear lime water, cork the bottle, and shake it. The lime water now becomes milky in appearance, and this milkiess proves the presence of carbonic acid gas. **Carbonic acid gas is a compound gas**, composed of the two elements carbon and oxygen, and hence it is sometimes called oxide of carbon, or, to distinguish it from another oxide of carbon, carbon dioxide. Thus it appears that, when carbon burns in oxygen, it combines chemically with the oxygen, generating heat and forming an entirely new compound. And further experiments in this direction would show that a certain weight of carbon always unites with the same corresponding weight of oxygen, producing an amount of carbonic acid gas equal in weight to the sum of the weights of its two constituents, and always generating a definite amount of heat, the quantity of heat being measured by the amount of water it will raise

through a certain number of degrees of temperature. It may also be noticed that the carbon taken out from the bottle after combustion (if it has not been entirely consumed) has lost in weight; but that which has disappeared has simply changed its form to assume the invisible gaseous state in combination with oxygen. Thus, 32 grains of oxygen gas are exactly sufficient for the complete combustion of 12 grains of carbon, and the burning of this quantity of carbon always produces exactly 44 grains of carbonic acid gas. **Combustion, then, is chemical combination, accompanied by heat and light.** The heat is produced by the chemical action—the **oxidation** of the carbon. We see also that when carbon burns there is no loss of substance; and in all cases of combustion, and, indeed, in all cases of chemical action, when a substance is apparently destroyed, it may be proved that it has simply changed its form, and that the change in form is never accompanied by any destruction of matter. Hence we say that **matter is indestructible.**

60. What takes place when a candle burns?

Experiment 33.—Place a dry bell-jar or an ordinary wide-mouthed bottle over a short lighted candle, and notice the rapid deposit of water on the inner surface of the glass. Notice, too, that after a short time the candle flickers feebly and finally goes out. This we know is due to the consumption of the oxygen gas originally contained in the air under the jar. Now raise the bell-jar slightly, and quickly slide a glass plate between it and the candle so as to shut in the air, which has evidently been changed in some way by the burning of the candle. Turn the bell-jar over, open it just a little, pour in some clear lime water, close again, and shake it. The lime water becomes milky, showing the presence of carbonic acid gas. This experiment proves that **when a candle burns, water and carbonic**

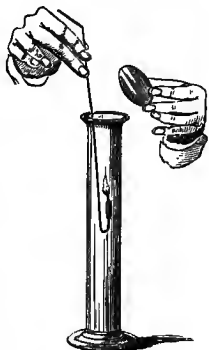


FIG. 20. — PASSING A LIGHTED CANDLE INTO A JAR OF OXYGEN.

acid gas are both produced. Here, then, is another example of combustion or oxidation, and another illustration of the indestructibility of matter. The candle contains the two elements carbon and hydrogen, and when it burns these two elements combine with the oxygen in the air, forming water (which is an oxide of hydrogen) and carbonic acid gas. The candle certainly decreases in weight, but its elements are not destroyed. If the candle be burnt in a jar of oxygen the same chemical changes will take place, but the oxidation will be more rapid, and therefore more heat and light will be generated *in a given time*.

61. Combustible substances contain carbon and hydrogen.

Most combustible substances contain the same two elements which were proved to exist in the tallow or wax of the candle. We may demonstrate this by repeating the experiments just mentioned, but using any other combustible substance, such as paper, wood, oil, gas, or spirit, instead of the candle. In every case we shall produce water and carbonic acid gas as the products of combustion, thus proving the presence of hydrogen and carbon.

62. The combustion of sulphur in oxygen.

Experiment 34.—Ignite a little sulphur (brimstone) in a small metal cup attached to a wire, and pass it into a jar of oxygen. It is seen to burn much more brightly in this gas than in air. As soon as the combustion is over, remove the sulphur, pour in a little water, close the jar, and shake it. Now pour in a little solution of litmus,¹ and notice that it immediately turns red. This change is explained thus. The sulphur, as it burned, combined with the oxygen, producing a suffocating gas called **sulphurous acid gas** or **sulphur dioxide**. This gas is soluble in water, and was therefore dissolved when shaken up with water in the jar. The presence of this new substance was proved by its action on the litmus solution. The product of combustion is the same when sulphur is burnt in air.

¹ Litmus is a vegetable colouring matter, naturally blue, but changed to red by the action of an acid.

63. The combustion of phosphorus in oxygen.

We have already noticed the changes produced by the burning of phosphorus in air (See Exp. 24), but it will be interesting to notice the increased energy of the combustion when the ignited phosphorus is passed into a jar of pure oxygen. The same white fumes of the oxide of phosphorus are produced, giving rise to the same acid substance when dissolved in water.

64. Iron will burn in oxygen gas.

Experiment 35.—Take a fine iron wire, and dip one end into flowers of sulphur. A little of the sulphur will probably adhere; if not, it may be made to do so by first heating the end of the wire. Now ignite the sulphur, and while it is burning pass it quickly into a jar of oxygen. The burning sulphur heats the iron wire to redness, and then the iron itself commences to burn brilliantly, continuing to do so till either the metal or the oxygen is consumed. On examining the jar we find numerous small particles of a black substance. This is an **oxide of iron**, commonly called the black oxide of iron. The weight of this oxide of iron is, of course, equal to the sum of the weights of the iron and oxygen consumed in its formation. Here we have an example of a substance which is combustible in pure oxygen, but not in the diluted oxygen of the atmosphere.

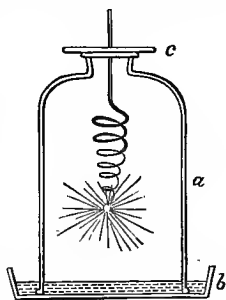


FIG. 21.—IRON WIRE BURNING IN OXYGEN.

a, Gas-jar, open top and bottom; *b*, tray of water to catch the falling particles; *c*, disc of wood or cardboard to hold the wire.

65. When a substance burns it increases in weight.

In each of the experiments just described it has been noticed that **bodies gain oxygen during combustion**. The following experiment will show that this oxidation causes an increase in the weight of the body burnt.

Experiment 36.—Balance some iron powder (*ferrum vidac-tum*) on a pan of a delicate pair of scales, and cause it to burn by applying heat. As the iron burns it combines with oxygen

of the air, and the increase in the weight of the iron will be proved by the falling of the pan containing it. The exact weight of oxygen necessary for the combustion of a given weight of iron may thus be determined ; and therefore the exact composition of the iron oxide produced will be known.

66. Summary.

The following is a summary of the main facts taught in the present lesson :—

1. Oxygen may be prepared by heating the red oxide of mercury.

2. The oxide of mercury is a compound, and is decomposed by heat into its two elements. Such a separation of a compound into its elements is termed analysis.

3. Oxygen may be most conveniently prepared by heating a mixture of potassium chlorate and manganese dioxide.

4. Substances which burn in air burn with greater brilliancy in pure oxygen.

5. When carbon burns in air or oxygen it produces carbonic acid gas.

6. Combustion is chemical combination accompanied by heat and light. When a substance burns it is not destroyed, but simply changes its form.

7. When a candle burns, water and carbonic acid gas are produced, proving that the candle contains both hydrogen and carbon.

8. All ordinary combustible substances contain these same two elements.

9. When sulphur burns, sulphurous acid gas is produced, and this gas, dissolved in water, produces an acid.

10. Phosphorus, when burnt in oxygen, produces white fumes of oxide of phosphorus.

11. Iron will burn in oxygen, the product of combustion being the black oxide of iron.

QUESTIONS ON LESSON V.

1. How would you prepare oxygen from the red oxide of mercury, and prove its presence?

2. Explain the term *analysis* as applied to chemistry, and give an illustration.

3. Describe fully some other method of preparing oxygen, and say how you would collect it in bottles or jars ; also sketch the apparatus you would use.
 4. What chemical change takes place when carbon is burnt in oxygen, and how would you prove that this change has taken place ?
 5. Explain fully what is meant by combustion.
 6. What do you mean when you say that matter is indestructible ? Describe an experiment illustrating your answer.
 7. What chemical changes take place when a candle burns ? Give experimental proofs.
 8. How would you prove the presence of carbonic acid gas ?
 9. What chemical elements are generally present in combustible substances ? Give proofs.
 10. What is the product formed by burning, (1) sulphur in oxygen, (2) sulphur in air ? How would you prove the presence of this product ?
 11. What substance is formed when phosphorus is burnt in oxygen ?
 12. How would you cause iron to burn in oxygen gas, and what is the product of combustion ?
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LESSON VI.

THE ACTION OF PLANTS AND ANIMALS ON THE AIR.

67. Carbonic acid gas is always present in the air.

Experiment 37.—We have already learnt how to detect the presence of carbonic acid gas (See Exp. 32), and by applying the lime-water test we may prove that this gas always exists in the atmosphere. For this purpose expose a shallow dish of lime-water to the open air for a few hours, and at the end of this time a thin solid film of carbonate of lime will be seen on the surface of the liquid. If, however, we expose some lime-water to the air of a room in which a number of persons are present, or in which gas jets or lamps are burning, the formation of the film will be much more rapid than in the open air, and this is especially the case if the room is not well ventilated. The greater abundance of carbonic acid gas in such a room is due to the combustion of the gas or oil, or, as we shall presently understand, to the respiration of the persons in the room.

68. How animals breathe.

Experiment 38.—We may now study the manner in which animals breathe, and how they affect the composition of the atmosphere. Pour some clear lime-water into a tumbler, and blow air into it from the lungs through a glass tube. In a very short time the lime-water becomes milky, proving that carbonic acid gas has passed into it. But whence this carbonic acid gas? It is formed in the body by the union of carbon with the oxygen of the air. The oxygen taken into the lungs passes through the thin walls of the air cells into the blood vessels surrounding them. By means of the blood system it is distributed to all parts of the body. Some of it, coming in contact with particles of carbon and hydrogen, combines with them, forming carbonic acid gas and water. We have already seen that when carbon and hydrogen burn they combine with oxygen, and form these same gases as the products of the combustion. We therefore say that **carbon and hydrogen are burnt up in the body; that the carbonic acid gas and water formed in the body, and afterwards exhaled, are products of the oxidation of carbon and hydrogen; and that the high temperature of the body is the result of this oxidation.** Thus we become acquainted with a permanent and abundant source of the carbonic acid gas in the air.

69. How plants breathe.

Plants also need air, but they do not require the same kind of nutriment as animals. The leaves of plants are sometimes called the respiratory organs, for it is chiefly by means of them that plants absorb the necessary gaseous food from the atmosphere. The following experiment will help us to understand how plants breathe, and the effect they produce on the composition of the air.

Experiment 39.—Place a growing plant or a freshly cut twig in a shallow dish of water, and cover it with a stoppered bell-jar. Now fill the jar with carbonic acid gas (to prepare carbonic acid gas see Exp. 62), replace the stopper, and

expose to sunlight for several hours. Carbonic acid gas is not a supporter of combustion—a lighted taper is instantly extinguished by it. Yet, if after the few hours' exposure to sunlight a lighted taper be passed into the jar, it will continue to burn, thus proving that much of the original gas has disappeared, and that oxygen gas has been exchanged for it. The plant has absorbed the carbonic acid gas, which, as we have seen, is composed of carbon and oxygen; it has retained the carbon and added it to its own solid structure, and returned the oxygen. **Thus we learn that plants require carbonic acid gas, and that they give off oxygen.**

70. Plants require both light and air.

If, instead of exposing the plant to sunlight, we place it in a dark cellar, this change will not take place. In fact, by an experiment similar to the one just described, it may be proved that **a plant absorbs oxygen and gives off carbonic acid gas during darkness.** This being the case, is it not unhealthy to keep plants in a room in which people live or sleep? Our experiments appear to prove that it is; but it must be remembered that although plants breathe like animals during the night, yet so little exchange of gases takes place during darkness that there would be no danger in sleeping in a conservatory with hundreds of plants. The breathing of plants takes place chiefly during the day time, and then the exchange of gases is exactly the opposite of that produced by animals.

71. The balance of life.

Those who have kept aquaria stocked with fish or other aquatic animals know how essential it is to keep up a supply of oxygen gas for the respiration of these animals. There are two methods of maintaining this supply of oxygen. One method consists in forcing a current of air into the water. Some of this air is dissolved, and the oxygen is afterwards separated as the water passes through the breathing organs of the animals. But the other method is by far the better. Growing plants are placed in the water. These plants not only supply the oxygen necessary for the animals, but they also

feed on the carbonic acid gas given off by these beings, and which, if allowed to accumulate in the water, would render it poisonous to animal life. Where the former method of aërating the water is employed, it is necessary to renew the water at intervals, in order to carry off the constantly increasing supply of carbonic acid gas ; but in the latter method, this very gas serves as a plant food, and the oxygen which it contains is given off by the vegetation and returned again to the animal in a pure state. It is possible to supply an aquarium with such a proportion of animals and plants, that the amount of oxygen gas produced by the latter is just sufficient to maintain a healthy condition of the former ; and so by keeping up a balance of animal and vegetable life, to render any renewal of the water altogether unnecessary. Such a balance of life is maintained on the earth ; and now we are able to explain how it is that, although carbonic acid is being continually poured into the air, yet there is no sensible increase in its amount.

72. Summary.

1. The atmosphere always contains a small amount of carbonic acid gas.

2. Animals cannot live without oxygen. They exhale carbonic acid gas.

3. The carbonic acid gas exhaled by animals is a product of the combustion of carbon in the body, and the high temperature of the blood in some animals is the result of this combustion.

4. Plants inhale carbonic acid gas, decompose it, retain the carbon, and exhale the oxygen.

5. The decomposition of carbonic acid gas in the plant takes place only under the influence of light.

6. The respiration of animals is exactly the opposite of that of plants, and the combined action of plants and animals maintains a practically unvarying amount of carbonic acid gas in the air.

QUESTIONS ON LESSON VI.

1. How would you prove the presence of carbonic acid gas in the air?
 2. Describe fully how animals breathe, and how their respiration affects the composition of the air.
 3. What is the cause of the high temperature of the blood in the higher animals?
 4. Describe fully the manner in which plants breathe, and the effect they produce on the composition of the atmosphere.
 5. What is the amount of carbonic acid gas in the atmosphere? Does this amount vary? Give reasons.
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LESSON VII.

THE THREE STATES OF WATER.

73. Water may exist as a solid, liquid, or gas.

WE are most familiar with the liquid form of water. But we know that when water is cooled down to a certain temperature, called its freezing point, it is converted into a crystalline solid which we call *ice*. Snow, hoar-frost, and hail are varieties of ice, the variation in structure being due to the varying circumstances under which the freezing takes place. We know also that when water is heated to a certain temperature, called its boiling point, it is converted into *steam*—the gaseous form of water. Although water is capable of existing in such a variety of forms, yet it never varies in chemical composition—ice, water, and steam are all composed of the same gases in exactly the same proportions.

74. Water expands when heated.

Experiment 40.—Procure a large glass flask with a narrow neck. Make a mark on the neck, and then pour in water till it stands exactly at this mark. Now heat it over a spirit lamp or gas burner, and notice that as long as the temperature of the water is rising, its volume continues to increase. Since the same weight of water now occupies a larger space, it follows that the heated water is lighter, volume for volume, than cold water. This may be proved also by carefully measuring out

equal volumes of hot and cold water in separate vessels, and weighing them on an ordinary pair of scales.

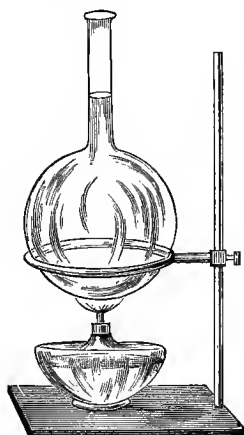


FIG 22.—ILLUSTRATING THE EXPANSION OF WATER.

75. Water does not expand and contract regularly.

Experiment 41.—As a rule substances expand while their temperature is rising, and continue to contract as long as they are cooling ; but in water we have an exception to this rule. For, if we fill a large flask having a very narrow neck with water, and then surround the flask with a freezing mixture of ice and salt, we shall see that the water continues to contract as it cools, only *till it reaches the temperature of 4°C .* (shown by placing a thermometer in the flask). The water at this temperature occupies its smallest volume, and is therefore at its

maximum density. If cooled still further it *expands* slowly till the freezing temperature is reached ; and, at this temperature, it expands considerably as it passes into the solid form called ice.

76. Water expands while in the act of freezing : therefore ice is lighter than water.

Experiment 42.—Take a piece of glass tubing, about 4 inches long, and $\frac{1}{4}$ -inch bore. Close one end in the blow-pipe flame, and draw out the other so as to leave only a very small opening. Now fill the tube with water as follows.—Stand the tube in an inclined position with its open end in water. Warm the air inside by means of a spirit-lamp. The air expands, and some bubbles pass out through the water. Allow the tube to cool ; the air contracts and water is forced in by atmospheric pressure, thus filling the space previously occupied by the expelled air. Remove the tube from the vessel of water, and boil the water in it till the steam has had time to drive

out all the enclosed air. Now invert the tube rapidly, and, while the water is still boiling, immerse the open end in cold water. The steam now condenses, and water is forced in till the tube is quite full. Seal the open end of the tube in the blow-pipe flame, and then place the tube in a mixture of ice and salt. After a few minutes the water freezes and the tube bursts, thus proving that expansion has taken place. One volume of water will produce about $1\frac{1}{10}$ volume of ice, both having exactly the same weight; therefore ice floats on water with about $\frac{1}{10}$ th of its volume above the surface. The specific gravity of ice is .916.

77. How water becomes heated.

Experiment 43.—Fill a large flask with water nearly up to the neck, drop into the water a small piece of some soluble colouring matter, such as aniline, and apply heat. The water at the bottom of the flask becomes heated first; and being lighter than the surrounding water, on account of its expansion, it rises. As fast as it ascends, other cooler water flows in to take its place, thus keeping up a constant circulation, and bringing every part of the water near the source of heat. This circulation of water is termed **convection**, and continues as long as the water is being raised in temperature. Similar movements are produced in the ocean by the heat of the sun, and in this manner **ocean currents** are formed.

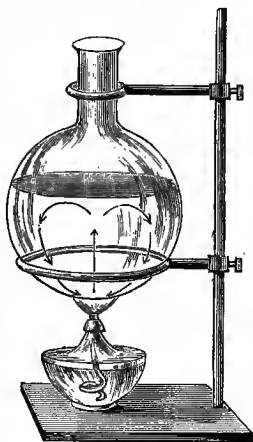


FIG. 23.—EXPERIMENT ILLUSTRATING CONVECTION.

78. The boiling point of water.

Experiment 44.—We have already seen that water is converted into a gas called steam when heated to a certain temperature. In order to study this change more carefully, heat

the water in a glass flask which contains also a thermometer. At first bubbles of air pass off, then after a time a sound is produced by the sudden formation and almost immediate disappearance of bubbles of steam at the bottom. This sound corresponds with the 'singing' of the kettle. The bubbles of steam formed at the hottest part of the vessel immediately come in contact with water at a lower temperature, and as steam cannot exist at a temperature lower than that of boiling water, they rapidly condense, *i.e.* pass back again to the liquid state. But after a time the whole of the water reaches a temperature of 100°C , and then the steam bubbles formed at the bottom pass through into the air, and the water is said to **boil**—the temperature has reached that point at which the vapour of the liquid is able to overcome the pressure of the atmosphere. Now watch the thermometer, and notice that **the temperature of the water ceases to rise after boiling commences**; and even if we increase the amount of heat, still there is no further rise. The additional heat imparted to the water passes off with the steam. And notice also that, while a dense cloud is passing from the mouth of the flask through the air, the flask itself is perfectly transparent. The flask contains steam, but steam is invisible; and when this invisible steam passes out into the cool air, it is condensed into a cloud composed of liquid particles of water. As the boiling point of water depends on atmospheric pressure, we should expect to find that water would boil at a lower temperature on a mountain than on low ground, and that down a mine below the sea-level the boiling point of water would be higher; this is actually the case. When water contains solid substances in solution it boils at a higher temperature; thus, sea-water boils at a higher temperature than fresh water.

79. Distillation.

Experiment 45.—When water passes off in the form of steam or vapour it leaves the greater part of its impurities behind. Therefore, if we boil water and condense the steam which passes off in a cold receiver, we shall obtain water almost entirely pure. This water is called **distilled water**, and the

apparatus used for the purpose, a **still**. A simple form of still consists of a glass flask for boiling the impure water, fitted with a delivery tube which passes into a similar flask used as a condenser, the latter being kept cool by a stream of cold water. A glass retort may be substituted for the boiling flask with advantage, as it obviates the necessity of fitting a cork, for the neck of the retort passes direct into the receiver. Fig. 25 represents **Liebig's condenser** with boiler attached; the principle of it will be easily understood from the sketch. When large

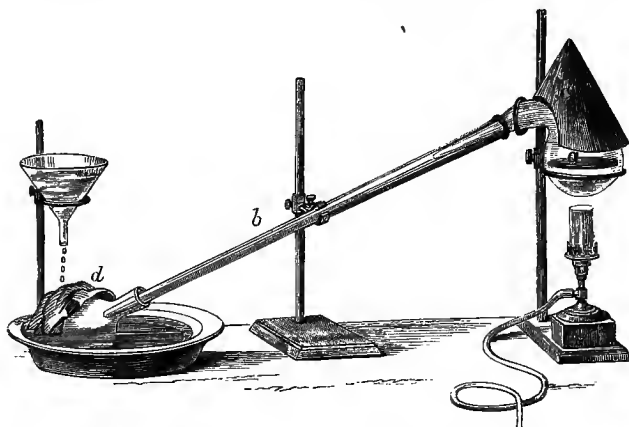


FIG. 24.—SIMPLE APPARATUS FOR DISTILLATION.

a, Retort, used as a boiler, and covered with a cone of tin plate to keep the upper part hot; *b*, glass tube communicating with the receiver or condenser; *d*, flask, used as a condenser, floating on cold water, and covered with a wet cloth, which is kept cool by a constant dripping of cold water from a funnel plugged with cotton-wool.

quantities of distilled water are required, the '**worm condenser**' will be found the most convenient form. This may be used in connection with an ordinary tea-kettle for a boiler, by joining the '**worm**' with the spout, and firmly closing the lid.

80. When water is converted into steam the volume increases enormously.

If we boil one pint of water till all has been converted into steam, and measure the volume of steam produced, we shall

find that the steam occupies about 1700 times the volume of the water from which it is formed, *i.e.* about 1700 pints. From this we may readily calculate that a quarter of a pint of water is sufficient to fill a hogshead with steam; and that a cubic inch of water will produce about a cubic foot of steam.

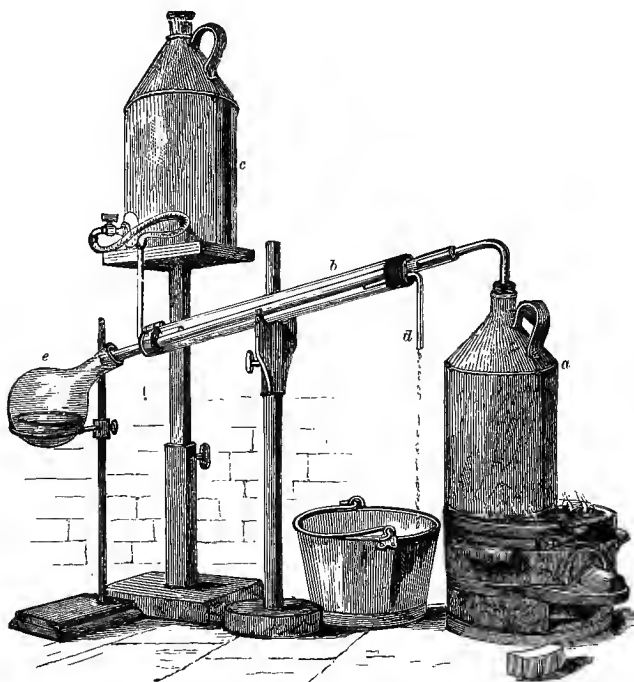


FIG. 25.—APPARATUS FOR DISTILLATION, INCLUDING LIEBIG'S CONDENSER.

a, An ordinary tin can, used as a boiler; *b*, Liebig's condenser, consisting of a large tube of cold water surrounding the central steam tube; *c*, supply of cold water to circulate through the outer tube of the condenser; *d*, exit for this water as it becomes warm; *e*, vessel to receive the distilled water.

Conversely, one hogshead of steam, when condensed by cooling, will produce only a quarter of a pint of water; and a cubic foot of steam will produce only one cubic inch of water.

81. Summary.

1. Water may exist as a solid, a liquid, or a gas; the difference in form being due to a difference in temperature.
2. Water expands when heated and contracts when cooled, although irregularly; and has its *maximum density* at 4°C .
3. Water expands when in the act of freezing, and ice is therefore lighter than the water from which it is produced. The specific gravity of ice is '916.

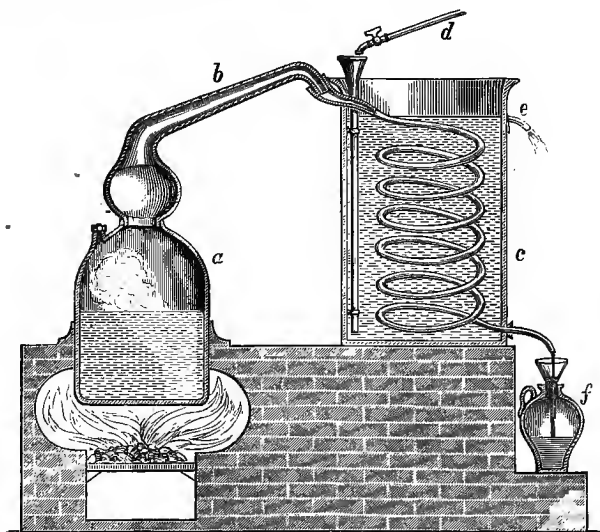


FIG. 26.—THE WORM CONDENSER.

a, boiler; *b*, steam tube; *c*, the worm condenser, consisting of a coiled tube surrounded by cold water; *d*, cold water supply; *e*, exit for water as it becomes warm; *f*, vessel to contain the distilled water.

4. Water is heated by *convection*—a circulation of its particles by which the heat absorbed at one point is distributed throughout its mass.
5. Water boils at 100°C , at which temperature it is converted into about 1700 times its volume of invisible steam.
6. When water boils, the steam is almost entirely free from the impurities that may be present. By condensing the steam we produce *distilled water*.

QUESTIONS ON LESSON VII.

1. Under what conditions is water changed into ice and steam?
2. Describe an experiment illustrating the expansion of water by heat.
3. What is meant by *maximum density* as applied to water? Illustrate your answer by an experiment.
4. Explain fully the meaning of the term *convection*.
5. What is the cause of the 'singing' of the kettle?
6. What is the boiling point of water, and how is it determined? Does water always boil at the same temperature?
7. What is meant by *distillation*? Describe some form of apparatus that may be used for the distillation of water.
8. Give a sketch of Liebig's condenser, and explain its use.
9. What volume of water is produced by the condensation of 100 gallons of steam?
10. How would you prove that water expands while freezing?
11. I have a solution of copper sulphate in water. How can I obtain the water in a pure state?
12. Describe an experiment by which you would ascertain the boiling point of alcohol.
13. I boil pure water, and then a solution of alum in water. Would there be any difference observable in the boiling points? If so, what?

LESSON VIII.

HYDROGEN.

2ⁿ 17804

82. The preparation of hydrogen gas.

Experiment 46.—Procure a wide-mouthed flask or bottle, and provide it with a sound and tight-fitting cork. Bore two holes in the cork; through one pass a thistle-head funnel, and through the other a straight glass tube. The tube of the funnel must pass to the bottom of the vessel when the cork is firmly set, but the other tube (the delivery tube) must simply pass through the cork. Now see that the apparatus is perfectly airtight as follows:—Close the top of the funnel by pressing the palm of the hand or a disc of india-rubber on it, and then suck out some of the air from the flask by means of the delivery tube. If now the tip of the tongue adheres firmly to the tube all is right. If not, the cork must be examined for the leakage, which may be stopped with paraffin wax; it is generally a good plan to soak the cork in melted paraffin before use. When

the apparatus is ready, remove the cork and put into the vessel about an ounce of zinc cuttings or zinc nails ; replace the cork, and pour through the funnel about two or three ounces of dilute sulphuric acid (one part of acid to about eight parts of water). Hydrogen gas begins to come off almost immediately, and this can escape only through the delivery tube, since the funnel dips into the liquid. Be careful that the apparatus is kept at a considerable distance from any flame. Instead of a wide-mouthed bottle for this experiment, a Woulffe's bottle may be used. More acid may be added occasionally as the formation of gas slackens.

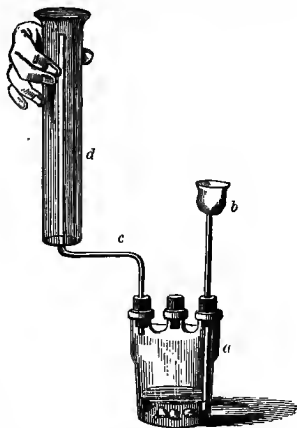


FIG. 27.—APPARATUS FOR PREPARING AND COLLECTING HYDROGEN.

a, Woulffe's bottle, containing acid and zinc ; *b*, thistle funnel ; *c*, delivery tube ; *d*, collecting jar.

83. The extreme lightness of hydrogen gas.

Hydrogen is the lightest gas known : the following experiments will illustrate its extreme lightness :—

Experiment 47.—Place a dry inverted bottle (or gas tube) over the delivery tube, and, after it has remained there for about a minute, transfer it to a flame, still keeping it in an inverted position. The gas will take fire with a sudden flash and a slight report ; and, if pure, will afterwards burn quietly for a few seconds with a pale blue flame. (Notice that the bottle is no longer dry, but is covered internally with a deposit of moisture.)

Experiment 48.—Fill the bottle again with gas in the same manner, but before transferring it to the flame, allow it to remain for a minute or two in the inverted position with its mouth open to the air. We again obtain the same result, viz. the combustion of the gas. These two experiments prove that the gas has risen into the bottle and displaced the air, and that it will remain in an open inverted vessel for some time,

and that it must, therefore, be lighter than air. This mode of collecting gas may also be used with all gases which are lighter than air, and is termed *collection by the upward displacement of air*. Hydrogen cannot be collected by *downward* displacement.

Experiment 49.—Fill the bottle again with hydrogen, and pass it into a second inverted bottle as shown in the figure.

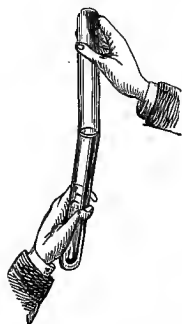


FIG. 28.—ILLUSTRATING THE MANNER OF PASSING HYDROGEN FROM ONE VESSEL INTO ANOTHER.

By applying the same test as before we can prove that the gas has passed from one bottle into the other. This also illustrates the natural tendency of hydrogen to rise in air from its lightness.

Experiment 50.—Suspend an inverted beaker or other thin glass vessel under one pan of a pair of scales, and balance carefully with weights in the other pan. Now allow hydrogen gas to pass into the beaker and displace the air, and the rising of the beaker will show that the hydrogen is lighter than the air displaced.

Experiment 51.—Connect an ordinary tobacco-pipe with the delivery tube of the hydrogen apparatus by means of about a foot of india-rubber tubing, and dip the bowl of the pipe for a moment into a solution of soap, so that a bubble may be blown by the gas. This bubble, when detached, will rise rapidly through the air. A small balloon of goldbeater's skin may also be made to ascend by filling it with hydrogen.

84. Water is formed when hydrogen is burned.

We have already seen that hydrogen is combustible, and that when it burns in a dry bottle a deposit of water globules is formed on the inner surface of the glass. We have also learnt (Lesson 5) that when a substance burns it combines with oxygen. This is the case with hydrogen; and water is a combination of this gas with oxygen. Hydrogen cannot burn unless in contact with oxygen, and when mixed with oxygen it burns so suddenly as to cause an explosion. Thus we noticed, on applying a light to an inverted bottle of hydrogen, that the

gas at the mouth of the jar, which was mixed with air, burnt suddenly with a slight report, while the purer gas burnt quietly, but only where it was in contact with air. It is on account of this explosive property of hydrogen when mixed with air that we must never apply a light to a stream of hydrogen issuing from the delivery tube until we are certain that all the air originally contained in the apparatus has been driven off. First collect some of the gas in a tube and test it; if it burns quietly and gradually, then it is safe to burn a jet at the delivery tube. Hold a dry glass vessel over a jet of burning hydrogen, and notice the deposit of water formed on the glass by the union of the gas with the oxygen of the air. Of course the weight of the water produced will be equal to the sum of the weights of hydrogen and oxygen entering into combination.

85. Hydrogen does not support combustion.

Experiment 52.—Fill a jar with hydrogen, and thrust into it a lighted taper. The gas burns at the mouth of the jar, but the flame of the taper is extinguished as soon as it passes from the air into the hydrogen. Hydrogen, then, although itself combustible, will not support the combustion of other substances.



FIG. 29.—SHOWING THE EFFECTS OF PASSING A LIGHTED TAPER INTO A JAR OF HYDROGEN.

86. Summary.

1. Hydrogen may be prepared by the action of sulphuric acid on zinc. (Hydrochloric acid may be substituted for the sulphuric, or iron for the zinc.)
2. Hydrogen is the lightest gas known (about $14\frac{1}{2}$ times lighter than air).
3. When hydrogen burns it combines with oxygen of the air and forms water; the weight of the water produced being equal to the sum of the weights of the hydrogen and oxygen that have combined.
4. Hydrogen gas does not support combustion.

QUESTIONS ON LESSON VIII.

1. How would you prepare hydrogen? Sketch the apparatus you would use.
2. Describe various experiments illustrating the extreme lightness of hydrogen gas.
3. How would you fill a bottle with hydrogen, and how may you know when the bottle is quite full?
4. What chemical change takes place when hydrogen burns?
5. What do you know of the chemical composition of water?
6. What occurs when a lighted taper is thrust into a jar of hydrogen gas?

LESSON IX.

THE CHEMICAL COMPOSITION OF PURE WATER.

87. How to determine the exact composition of water.

Experiment 53.—Procure a strong glass tube, about 18 ins. long, closed at one end, and graduated accurately for the measurement of the volume of its contents. This tube must also be provided with two platinum wires fused through the glass near the closed end and nearly touching each other inside. Such a tube is called a **eudiometer**, or water-measure. Fill this tube with mercury, close with the thumb, invert, and place it upright in a vessel of mercury. The thumb may now be removed, and the pressure of the atmosphere will support the column of mercury in the tube. Pass into the tube equal volumes of hydrogen and oxygen, each gas being carefully measured, by means of the graduated scale, as it displaces the mercury. Now cause an electric spark to pass between the ends of the platinum wires by connecting the wires with a suitable galvanic apparatus. This ignites the mixture, part of which is suddenly burnt, forming a deposit of water inside the tube. The mer-

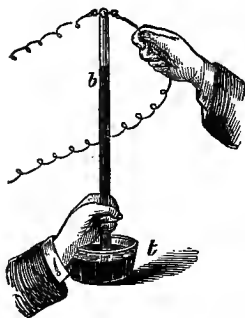


FIG. 30.—ILLUSTRATING THE MANNER OF USING THE EU-DIOMETER.

b. The eudiometer tube; *t*, a vessel containing mercury.

cury now rises in the tube till only one fourth the volume of gas remains. The remaining three-fourths of the gas has been consumed, giving rise to the deposit of water which was seen on the tube. Now remove the tube and its contents from the vessel of mercury, and turn it over so as to bring the gas to the open end. Put a lighted taper in—it burns brilliantly. The gas, therefore, is oxygen. Hence we learn that the whole of the hydrogen and half of the oxygen have entered into combination to form water, and that the other half of the oxygen remains unchanged. If we try this experiment over and over again, varying the proportion of gases in the mixture, we always obtain the same result—viz. a given volume of hydrogen combines with *exactly* half its volume of oxygen, forming water. If we mix the gases in these proportions exactly, nothing will remain after combustion but the small drop of water formed; and, if either gas be in excess, the excess will remain unchanged. Thus we learn that water is composed of two volumes of hydrogen and one volume of oxygen.

Experiment 54.—This may be further demonstrated by passing a current of electricity through water slightly acidulated. The electricity will decompose the water into its two gases; and if the two gases be collected in separate tubes, the hydrogen will always be found to have double the volume of the oxygen.

Experiment 55.—If we mix hydrogen and oxygen in a soda-water bottle (this kind of bottle being selected on account of its strength) and apply heat to the mixture, the two gases combine with a loud explosion, forming water. We may observe that the force of the explosion is greatest when we mix the gases in the proportions in which they combine to form water.

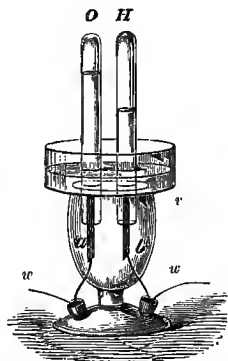


FIG. 31.—APPARATUS FOR DECOMPOSING WATER BY MEANS OF THE ELECTRIC CURRENT.

w, wires connecting the apparatus with a galvanic battery; *a* and *b*, strips of platinum at the ends of these wires, one passing into each of the tubes O and H; *v*, a vessel containing the water to be decomposed; O and H, the tubes for collecting the oxygen and hydrogen.

88. The composition of water by weight.

Experiment 56.—The exact composition of water by weight may be determined by passing steam through an iron tube containing red-hot iron borings. Before doing this, however, the exact weight of the iron tube and its contents must be noted. As the steam passes over the red-hot iron it is decomposed, its oxygen combining with the iron and forming an oxide of iron, while the hydrogen thus set free passes over, and may be collected in a vessel standing over water (thus proving it is not steam, which would mingle with the water). After this chemical change has been going on for a short time, allow the apparatus

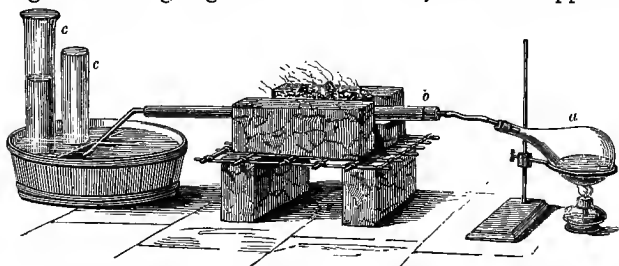


FIG. 32.—APPARATUS FOR PASSING STEAM OVER RED-HOT IRON AND COLLECTING THE HYDROGEN PRODUCED.

a, flask containing water; *b*, tube containing red-hot iron; *c*, jars for collecting the hydrogen formed.

to cool, and again weigh the iron tube. It will be noticed that there is an increase in weight, this increase showing the weight of oxygen which has combined with the iron. This, together with the hydrogen collected, will of course weigh the same as the steam decomposed. Let us suppose, for instance, that the tube contains 42 grains of iron borings. This will be found to be exactly sufficient for the decomposition of 18 grains of water in the form of steam. Exactly 16 grains of this are oxygen, and combine with the iron, and 2 grains of hydrogen pass over into the collecting vessel. Such an experiment as this proves to us that water contains one-ninth its own weight of hydrogen (2 parts in 18) and eight-ninths its weight of oxygen. We have already seen that the composition of water by *volume* is 2 parts of hydrogen to one of oxygen, and the difference between

this and the composition by weight will be readily explained when we remember that oxygen is 16 times heavier than hydrogen.

89. Summary.

1. Water contains two volumes of hydrogen to one of oxygen.

2. When a mixture of hydrogen and oxygen is exploded, the gases always combine in the above proportion, any excess of either gas remaining unchanged.

3. When water is decomposed by means of the electric current, it always yields these gases in this same proportion.

4. Steam may be decomposed by passing it over red-hot iron, the oxygen combining with the iron and the hydrogen being set free.

5. The increase in the weight of the iron by this combination is equal to eight-ninths of the weight of the steam decomposed. The remaining one-ninth of the steam is hydrogen.

QUESTIONS ON LESSON IX.

1. Describe the eudiometer and its use.

2. What is the result of firing a mixture of four volumes of hydrogen and six volumes of oxygen in the eudiometer?

3. Ten volumes of hydrogen are mixed with four volumes of oxygen, and the mixture exploded. What volume of gas remains, and how will you prove which gas it is?

4. How would you prove the composition of water by means of the electric current? Make a sketch of the apparatus you would use.

5. How would you prepare hydrogen by means of water and iron? Sketch the apparatus necessary for this experiment.

6. Describe fully an experiment by which the composition of water by weight may be determined.

7. How much hydrogen may be obtained by the decomposition of 600 grains of steam?

8. What is the percentage composition of water by weight?

9. If 20 grains of steam be passed over red-hot iron, what will be the increase in the weight of the iron?

10. What weight of oxygen is necessary for the complete combustion of 9 grains of hydrogen?

11. Six pints of oxygen are to be converted into water by means of a jet of burning hydrogen. What volume of hydrogen will be necessary for this?

LESSON X.

CARBON.

90. The varieties of carbon.

CARBON is an elementary substance. There are several varieties of carbon, differing so much from each other in their properties, that we could not regard them as being chemically the same substance without a clear proof that such is the case. These varieties are :—

(1) **Charcoal**, prepared from animal and vegetable substances by heating them while excluded from the air.

(2) **Blacklead** (also called **graphite** and **plumbago**), a mineral substance, obtained largely in Cumberland and other parts.

(3) **Lampblack**, a variety of soot ; a deposit of finely divided carbon formed by the incomplete combustion of substances containing that element.

(4) **Diamond**, a crystalline variety of carbon, particularly hard.

91. The preparation of charcoal.

All animal and vegetable substances contain carbon, and when they are heated while in any way excluded from the air

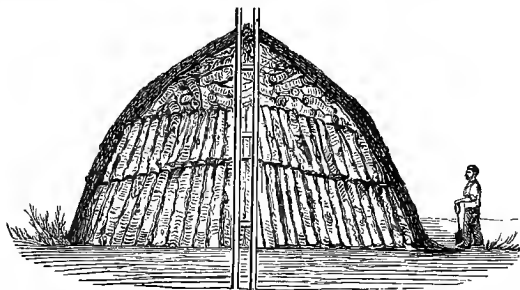


FIG. 33.—ILLUSTRATING THE PREPARATION OF CHARCOAL ON A LARGE SCALE.

(as in closed vessels, for instance), black masses of impure carbon are formed, called **vegetable charcoal** when prepared

from wood or other parts of plants, and **animal charcoal** when obtained from animal substances. **Vegetable charcoal** is prepared in large quantities by causing a pile of blocks of wood to smoulder while covered over with a layer of turf. During this slow and imperfect combustion of the wood, various gases and vapours are given off, and these are allowed to escape through an opening at the top of the pile.

Experiment 57.—This process may be imitated, on a smaller scale, by heating sawdust in a test-tube over a spirit-lamp flame. Various inflammable gases and vapours which are driven off may be ignited at the top of the tube, and as all these proceed from the sawdust, it will be found that the charcoal formed weighs much less than the wood from which it is prepared. Similar results may be obtained by heating dried meat, bread, starch, &c. We may prove that sugar contains carbon by heating it strongly on a metal plate, for when so treated a mass of charcoal is formed.

92. The combustion of carbon.

In Lesson V. we learnt that when charcoal is made to burn in oxygen it combines with the oxygen, forming a gas called carbonic acid gas, which turns lime-water milky. The same result is obtained when any form of carbon, or any organic substance, is burnt in oxygen or air.

Experiment 58.—Place a bell-jar over a lighted candle till all the oxygen is consumed, and the flame is extinguished. Then quickly raise the jar and cover its mouth with a glass disc, so as to retain as much as possible of the gas produced by the combustion. Pour in a little clear lime-water, and shake it. The lime-water immediately becomes milky, thus proving the presence of carbonic acid gas, and also the existence of carbon in the candle. This experiment may be repeated with various combustible substances, such as paper, wood, oils, spirit, &c.

93. Carbon does not combine with oxygen at ordinary temperatures.

Experiment 59.—Pass a current of pure air (not air from the lungs), or oxygen, through a hard glass tube containing

small pieces of charcoal, and then allow this air to pass into lime-water. No change takes place. But now apply heat to the charcoal, still keeping up a slow current of air. When the temperature has reached a certain point, the charcoal

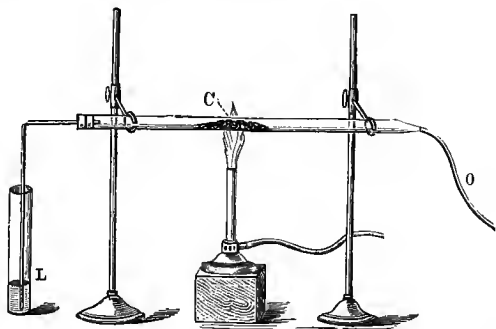


FIG. 34.—APPARATUS FOR ILLUSTRATING THE COMBUSTION OF CHARCOAL.

o, Tube conveying air or oxygen; *C*, hard glass tube containing charcoal; *L*, vessel containing lime-water.

begins to glow, and immediately after this a milkiness is produced in the lime-water. Thus we see that a certain amount of heat is necessary to make carbon combine with oxygen.

94. Charcoal is not pure carbon.

Continue the above experiment till the carbon ceases to burn. Allow the tube to cool, and, when cold, shake out its contents. These will consist of a whitish powder, weighing much less than the charcoal with which we commenced the experiment. The charcoal consisted chiefly of carbon, and the whole of this combined with oxygen and passed off as carbonic acid gas. That which remains is called the ash, and is composed of certain mineral substances. This mineral (inorganic) matter existed in the wood from which the charcoal was prepared, and, therefore, formed part of the tree which supplied this wood. Nearly all animal and vegetable substances contain a small amount of mineral matter, and in some cases a large proportion is present. The amount of mineral substance may always be determined by burning. The **organic** portion passes

off as carbonic acid and other gases, while the **inorganic**, or mineral portion, remains behind as the ash. Acids and alkalis have no effect on charcoal.

95. How to prove that the diamond is carbon.

When a diamond is burnt in oxygen, it always forms carbonic acid gas. This proves that the diamond *contains* carbon. But does it contain anything else? It does not ; for carbonic acid gas is the only gas produced by its combustion, and little or no ash remains. Similarly it may be proved that graphite is also a variety of carbon—an elementary substance, and *not a compound substance containing carbon*.

96. Summary.

1. Charcoal, blacklead, lampblack, and the diamond are varieties of the element carbon.

2. All these produce carbonic acid gas when burnt.

3. Charcoal is prepared by heating organic matter while excluded from the air, during which process various inflammable gases and vapours are given off.

4. Carbonic acid gas, the product of the combustion of carbon, will turn lime-water milky ; hence lime-water is used to prove the presence or absence of this gas.

5. Carbon will not combine with oxygen unless heated strongly.

6. Charcoal, when completely burnt, leaves an ash composed of inorganic or mineral substances. Most organic substances leave a similar ash.

7. Carbon is an essential constituent of all organic substances.

QUESTIONS ON LESSON X.

1. Give the names of the varieties of carbon, and a short description of each.

2. How is charcoal prepared on a large scale?

3. How would you prepare a small quantity of charcoal? Describe any changes that occur during the preparation.

4. Prove that white sugar contains carbon.

5. What chemical change takes place when charcoal is burned? How would you prove that this change takes place?

6. Give a list of several substances containing carbon.

7. How would you prove that carbon will not combine with oxygen unless heated?

8. Does charcoal contain anything beside carbon? If so, what? and how would you obtain it?

9. A piece of wood is completely burnt. What has passed off, and what remains?

10. How would you prove that the diamond is carbon?

LESSON XI.

CHALK, LIME, AND CARBONIC ACID GAS.

97. Structure and composition of chalk.

CHALK is a rock composed of marine shells, most of them so small that we can discern them only with the help of a microscope. It is a compound substance containing three elements. These are **calcium**, **carbon**, and **oxygen**. The properties of the last two of these elements have already been described. Calcium is a yellow metal which readily combines with oxygen when exposed to the air. In fact, its affinity for oxygen gas is so powerful that it is very difficult to obtain the metal in a pure state. Chalk is a variety of **carbonate of lime**.¹ A **carbonate** is a substance formed by the union of carbonic acid gas with the oxide of a metal. Chalk, then, is a compound composed of two simpler compounds, viz. calcium oxide and carbonic acid gas. Lime is the common name for the oxide of the metal calcium. The following table will show clearly the exact composition of chalk by weight; and the object of the present lesson is to show how this may be determined, as well as to demonstrate the properties of the two binary compounds which enter into the composition of chalk.

100 parts, by weight, of chalk contain

44 parts of carbonic acid gas, consisting of	{	Carbon . . .	12
		Oxygen . . .	32
and 56 parts of lime, containing . . .	{	Calcium . . .	40
		Oxygen . . .	16
			<u>100</u>

¹ Sometimes termed *calcium carbonate*.

Hence the percentage composition of chalk is

Calcium	40
Carbon	12
Oxygen	48

98. The preparation of lime from chalk.

When chalk is strongly heated it gives off its carbonic acid gas and only the lime remains. If a piece of chalk be heated to redness in an open crucible, it is reduced to lime with a loss of weight. Lime is sometimes prepared in large quantities by heating chalk in kilns, coal being mixed with the chalk and burnt with it. In this way 100 tons of chalk may be made to yield *about* fifty-six tons of lime. *Theoretically* it should yield *exactly* fifty-six tons; but we must allow in practice for the impurities present in both chalk and lime, and for the unavoidable waste through crumbling, etc. There are **other varieties of carbonate of lime** which may be used for the preparation of lime. These are **marbles, limestones**, and various **shells**. Although these differ so much in general appearance, they have all the same composition, and produce exactly the same results when heated.

99. The properties of lime. Lime-water.

Pure lime is a white infusible solid. When intensely heated it glows very brilliantly, producing the light known as the lime-light, but does not suffer any decrease in weight, for it does not burn. Unlike limestone, lime combines very readily with water, generating great heat, and forming a white powder called **slaked lime**. Fifty-six parts, by weight, of lime are just sufficient to combine with eighteen parts of water. If we add the water to the lime in this proportion, the water forms with it a perfectly dry powder. When lime is mixed with excess of water, it forms a paste which will gradually set. On this account it is used for making **mortar**, which is a mixture of lime and sand with water, together with some binding material such as hair. **Slaked lime** (sometimes called *calcium hydrate*) is slightly soluble in water, and its solution is called **lime-water**.

Experiment 60.—Lime-water may be conveniently prepared by adding a small quantity of lime or slaked lime (about

half an ounce) to a gallon of water, stirring well, and then, after the whole has remained at rest for several hours to allow the excess of solid matter to settle at the bottom, decanting the clear liquid.

100. Other compounds containing lime.

In addition to the various limestones, marbles, and shells, there are many other substances containing lime. Among these may be mentioned the rocks **gypsum** and **alabaster**; both these are varieties of **sulphate of lime**. They consist of lime, sulphur, and oxygen, together with water. When gypsum is heated till the water has been driven off, and afterwards powdered, **plaster of Paris** is obtained. This, when mixed with water, combines again with the liquid and *sets* to a solid mass. This property renders it valuable for making moulds and casts. The bones of animals contain much **phosphate of lime**—lime in combination with phosphorus and oxygen.

Experiment 61.—By steeping a bone in dilute hydrochloric acid for some time, all this mineral phosphate may be dissolved out, leaving only the organic portion of the bone in the form of a flexible mass of gristle. On the other hand, by burning a bone, we destroy the animal matter, and a white ash is left, consisting chiefly of phosphate of lime.

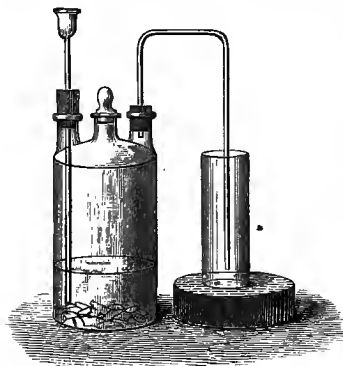


FIG. 35.—APPARATUS FOR THE PREPARATION AND COLLECTION OF CARBONIC ACID GAS.

101. The preparation of carbonic acid gas.

Experiment 62.—It has already been seen that carbonic acid gas may be obtained from chalk, limestones, and marbles by the application of heat. But this gas may be more conveniently prepared by adding an acid to either of these substances, or to any other carbonate; for all carbonates contain carbonic acid gas. Common **washing soda** is really

a carbonate of soda, and answers the purpose well. **Sherbet** contains bicarbonate of soda and tartaric acid. Both these substances are solids, and two solids cannot give rise to chemical action ; but, on the addition of water, the acid decomposes the carbonate, and sets free carbonic acid gas. Vinegar (an impure variety of acetic acid) will set free the same gas from egg-shell or any other carbonate. Perhaps the best method of preparing carbonic acid gas is by adding hydrochloric acid to chalk or marble. In this case, the chlorine of the acid combines with the metal calcium, forming chloride of calcium, which may be obtained in the solid state by the evaporation of the liquid remaining after the gas has ceased to come off.

102. Properties of carbonic acid gas.

This is an invisible gas ; and, being heavier than air, it may be collected by the downward displacement of air as shown in fig. 35. Several jars should be filled with the gas for the illustration of these and other properties. To ascertain when each jar is full, pass into it a lighted taper ; the flame will be extinguished the moment it passes into the gas: and, at the same time, it will be noticed that the gas itself is not combustible.

Experiment 63.

Pour carbonic acid gas from one vessel into another as an illustration of its density ; and test the result with a lighted taper.

Experiment 64.

Procure a vessel, provided with a tap.

Fill it with carbonic acid gas, draw off some of the gas by turning the tap, and test the result as before.

Experiment 65.—Balance an open glass vessel on a pair of

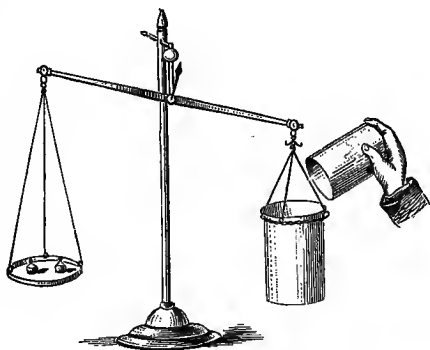


FIG. 36.—ILLUSTRATING THE DENSITY OF CARBONIC ACID GAS.

scales ; fill it with carbonic acid gas, and notice how the weight increases as the air is being displaced by this gas.

Experiment 66.—Fill a wide vessel with carbonic acid gas. Now blow soap bubbles with air, let them fall into this vessel, and they will float on the gas.

Experiment 67.—Allow some carbonic acid gas to bubble through water ; some of it will remain dissolved in the water, and the presence of this gas may be proved as follows :—

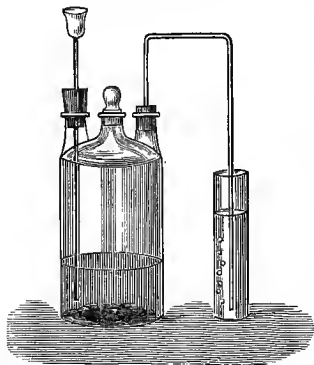


FIG. 37.—SHOWING HOW A SOLUTION OF CARBONIC ACID GAS MAY BE PREPARED.

(a) Add a little clear lime-water to a portion of the solution, and a milky appearance will immediately be produced. This is caused by the conversion of lime into chalk.

(b) Add a blue solution of litmus to another portion, and the litmus will slowly turn red, thus proving that the solution of carbonic acid gas has weak acid properties.

(c) Apply heat to a third portion of the solution of carbonic acid gas. The gas will soon be driven off, and may be collected and again tested with lime-water.

103. The formation of calcium carbonate by synthesis.

The lime-water test for carbonic acid gas has been frequently mentioned, but has not as yet been fully explained. When carbonic acid gas is passed into lime-water, the gas combines with the lime in the water, forming calcium carbonate ; and it is the formation of this insoluble carbonate which renders the water milky. The two binary compounds—lime and carbonic acid gas—combine to form a more complex compound—calcium carbonate. This is an illustration of chemical **synthesis**, or building up. By the action of acid on 100 grains of pure chalk or marble we can produce 44 grains of carbonic acid gas ; and if the whole of this gas be passed into sufficient

lime-water, it may be made to combine with 56 grains of lime, thus producing again 100 grains of carbonate of lime.

104. Calcium carbonate is soluble in water containing carbonic acid gas.

Experiment 68.—Pass carbonic acid gas into lime-water for some considerable time. At first the gas combines with the lime as above stated, forming the insoluble carbonate of lime. But after the whole of the lime present has thus been converted into the carbonate, the excess of the gas dissolved in the water begins to convert the insoluble carbonate into a soluble *bicarbonate*, and after a short time all the solid matter is dissolved, the liquid becoming quite clear. By boiling this clear liquid in a test tube or flask, the excess of carbonic acid gas may be driven off; and the bicarbonate being again converted into the insoluble carbonate, the liquid again becomes milky, and some of the solid matter is deposited on the sides of the vessel. The water of rivers and springs always contains carbonic acid gas, and is therefore capable of holding carbonate of lime in solution. Such water, when boiled, parts with its carbonic acid gas, at the same time depositing its lime. In this way our kettles and boilers become covered with an incrustation called '*fur*.' This *fur*, being a carbonate, will yield carbonic acid gas when treated with an acid. Many limestone rocks have been formed from other limestones by solution and subsequent deposition brought about in a similar manner.

105. Summary.

1. Chalk is a variety of carbonate of lime, containing the elements calcium, carbon, and oxygen. Its percentage composition being calcium, 40; carbon, 12; oxygen, 48.
2. All limestones and marbles have the same chemical composition as chalk.
3. When any form of calcium carbonate is strongly heated, carbonic acid gas is given off and lime remains.
4. One hundred parts, by weight, of the carbonate will yield 56 parts of lime, and give off 44 parts of carbonic acid gas.

5. Lime is the oxide of the metal calcium.
6. Lime combines energetically with water, forming slaked lime or calcium hydrate.
7. Lime-water is a solution of slaked lime in water.
8. Gypsum and alabaster are forms of the sulphate of lime, and consist of lime, sulphur, and oxygen ; they also contain a certain amount of water.
9. When gypsum is heated to drive off this water, and afterwards powdered, plaster of Paris is produced.
10. The bones of animals contain large quantities of phosphate of lime. This may be dissolved out by an acid.
11. Carbonic acid gas may be prepared by the action of any acid on any carbonate.
12. Washing soda is carbonate of soda.
13. Carbonic acid gas :—
 - (a) Is invisible.
 - (b) Is heavier than air ($22 : 14\frac{1}{2}$).
 - (c) Is not combustible.
 - (d) Does not support combustion or animal life. (But is the great supporter of vegetable life.)
 - (e) Is soluble in water.
 - (f) In solution possesses weak acid properties.
 - (g) Forms calcium carbonate with lime-water.
14. Water containing carbonic acid gas has the power of dissolving small quantities of calcium carbonate. The gas may be driven off by heat, and then the carbonate again returns to its insoluble form.
15. The '*fur*' deposited on kettles and boilers is a carbonate of lime, formed by the expulsion of the dissolved carbonic acid gas which enabled the water to dissolve the carbonate.

QUESTIONS ON LESSON XI.

1. What are the chemical elements entering into the composition of chalk ? In what proportions are these elements present ?
2. Describe each of the elementary substances present in a piece of chalk.
3. What is a carbonate ? Give a few examples of carbonates.
4. Give the exact compositions of lime and carbonic acid gas.
5. How is lime prepared ?

6. Name some varieties of carbonate of lime.
7. How much lime may be obtained from 50 lbs. of carbonate of lime? How much carbonic acid gas would be liberated during the process of preparation?
8. What do you know of the properties of lime?
9. What is lime water? How would you prepare it?
10. What is slaked lime? By what other name is it known?
11. How much water is needed for the 'slaking' of one hundredweight of lime?
12. What do you know of gypsum, alabaster, and plaster of Paris?
13. How would you separate the animal from the mineral matter in a bone? Of what does this mineral matter chiefly consist?
14. What do you know of the chemical composition of common washing soda?
15. Describe a convenient method of preparing carbonic acid gas. Sketch the apparatus you would use.
16. Enumerate the properties of carbonic acid gas, and describe the experiments by which you would illustrate these properties.
17. How would you prepare calcium carbonate? What would be the exact quantities of materials necessary for the formation of 500 grains of the carbonate?
18. Explain what is meant by chemical *synthesis*. Give an illustration.
19. What result is produced when carbonic acid gas is passed through lime water for a considerable time?
20. Describe fully the formation of the 'fur' in a boiler.
21. What are the binary compounds present in a piece of limestone? How would you separate these binary compounds, and how would you cause them to recombine?

LESSON XII.

COAL—COAL-GAS—FLAME.

106. Manufacture of coal-gas.

Coal-gas consists chiefly of a compound gas containing the elements carbon and hydrogen. It is obtained by the distillation of coal; but as a large number of vapours distil over, many of which interfere with the lighting properties of the coal gas, it has to be purified in various ways before it is supplied to us for illuminating purposes. The **process of manufacture** may be illustrated in a simple manner as follows:—

Experiment 69.—Heat some coal-dust in a small flask made of hard glass and provided with a cork and delivery tube. In a short time a mixture of various vapours will begin to pass off,

together with a certain amount of smoke. Some of these vapours will condense on the upper part of the flask and in the delivery tube. The presence of these vapours will interfere

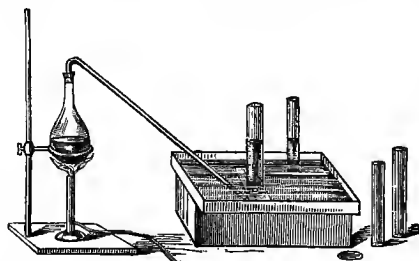


FIG. 38.—ILLUSTRATING A SIMPLE METHOD OF PREPARING AND COLLECTING COAL-GAS FROM COAL-DUST.

with the combustion of the coal-gas to such an extent that it will be difficult to obtain a steady flame, but if collected over water, many of the impurities will be separated, and the clear gas thus obtained will burn almost as well as that supplied to

our towns. Let the application of heat be continued till the gas ceases to be evolved, and then, on breaking the flask, we obtain a porous mass of **coke**, which necessarily weighs less than the coal used. In gas-works the coal is heated in large iron retorts; and the coke, which is removed from these retorts while still red hot, is immediately swamped with water, and afterwards broken and sold for fuel.

107. Properties of coal-gas.

Coal-gas contains much hydrogen, and is very similar to hydrogen in its properties; so much so that it is often used as a substitute for that gas, especially as it is generally conveniently at hand, and is much less expensive. Coal-gas, like hydrogen, is—

1. Combustible.
2. Lighter than air.
3. Explosive when mixed with air or oxygen.
4. Not a supporter of combustion.

It is this gas that is generally used for filling balloons, for although it is not so light as hydrogen, and consequently a larger volume would be required to lift the same weight, yet it is the less expensive of the two. Unlike hydrogen, coal-gas burns with a bright flame, the hydrogen, as will be remembered,

burning with a pale blue flame. The above-mentioned properties of coal-gas may be illustrated by a few simple experiments, such as those described in the lesson on hydrogen.

108. Combustion and composition of coal-gas.

Experiment 70.—Hold a dry glass vessel over the flame of burning coal-gas, and observe the deposit of water formed on the surface of the glass. The formation of this water is a proof that coal-gas contains hydrogen. Collect some of the gases produced by the combustion of coal-gas in an inverted gas-jar, and shake it up with a little lime-water. Chalk will be formed, thus proving the presence of carbonic acid gas, which must have been produced by the combustion of carbon, since unburnt coal-gas will not form chalk with lime-water. Thus we learn that **coal-gas contains both carbon and hydrogen**. Coal-gas consists chiefly of two distinct compound gases, both containing carbon and hydrogen. They are—

1. **Heavy carburetted hydrogen**, or **olefiant gas**, and
2. **Light carburetted hydrogen**, or **marsh gas**.

The former is the principal illuminating constituent. The latter is called marsh gas because it is produced by vegetable decay, especially in marshy districts; it is also called **fire-damp** by the miner, because, when mixed with air and coming in contact with a light, it is the cause of disastrous explosions in mines.

109. Various other substances obtained from coal.

It has already been stated that a variety of substances are produced by the distillation of coal. Many of these products are of great commercial value, as they are useful in various arts. The chief of these (excluding the coal-gas and coke already mentioned) are *tar*, *pitch*, *ammonia*, *benzine*, *naphtha*, *carbolic acid*, and *aniline*. The **coal-tar** is deposited, with other substances, in a condenser which is kept cool by currents of air. When this tar is distilled, a large number of substances come off, and **pitch** is left behind. The beautiful **aniline colours** are produced by the distillation of tar. A substance called *ammoniacal liquor* is deposited with the tar in the condenser, and from this, by the aid of sulphuric acid, large supplies of

ammonium sulphate are obtained. In fact, this ammoniacal liquor is the chief source from which we obtain the compounds of ammonium.

110. Flame.

The flame of a candle is really the flame of burning *gases*, for the wax or tallow of the candle is reduced to a gaseous state by heat before its combustion actually commences. This is also true as regards oil, spirit, and other ordinary combustibles ; and the remarks on flame in the present article apply equally to the flames produced by burning any of these substances.

Experiment 71.—Lower a piece of wire gauze over a candle flame from above, till it is brought to the level of the middle of the flame. It will now be seen that the gauze is heated in such a manner as to produce a luminous ring, surrounding a dark spot. This proves that the flame is not solid, but is rather a hollow cone. The central portion contains the unburnt carbon and hydrogen, which cannot burn till they come in contact with the oxygen of the air. Now remove the gauze, and notice that the true flame consists of two distinct portions—a luminous part, which surrounds the dark inner zone ; and a transparent and almost non-luminous zone, of a pale blue colour, surrounding the lower portion of the flame.

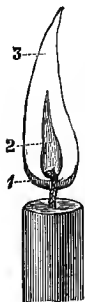


FIG. 39. — THE CANDLE FLAME.

1, pale blue zone ;
2, hollow core ; 3,
luminous zone.

The candle flame, then, consists of three distinct parts :—

- (1) Inner dark core of unburnt gases.
- (2) Luminous zone, giving brilliant white light.
- (3) Pale blue zone.

Experiment 72.—The hollowness of a flame may also be illustrated by lowering into it a piece of stout white paper, and keeping it still for a second or two. The paper will then exhibit a brown ring, produced by scorching, surrounding a clean circle of unscorched paper. A still more interesting illustration may be seen in the following experiment :—

Experiment 73.—Bend a piece of glass tube as shown in

the accompanying figure, and pass one end into the inner zone or hollow of a candle flame. A little of the unburnt gas forming the inner zone will pass through the tube, and may be ignited at the other extremity.

The brilliant white light of the flame is caused by numerous white-hot particles of unburnt carbon. The pale blue portion of the flame is due to the combustion of hydrogen.

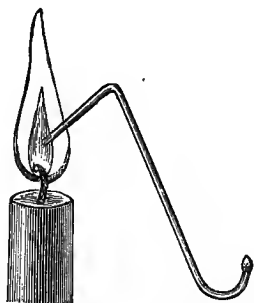


FIG. 40.—ILLUSTRATING THE HOLLOWNESS OF A CANDLE FLAME.

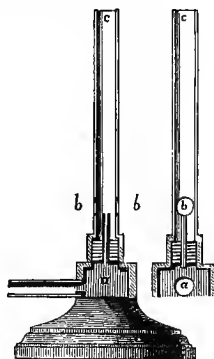


FIG. 41.—THE BUNSEN BURNER
a, gas jet; *c*, tube surrounding the jet; *b* marks the positions of the openings in the tube for the admission of air.

111. The Bunsen burner.

The **Bunsen burner** is a burner in which air is mixed with the coal-gas previous to combustion. The coal-gas is admitted through a tube at the bottom, and immediately mixes with air, which enters through two holes. The result on ignition is a pale blue flame, which gives greater heat, and which will not cause a deposit of carbon on anything held in it.

A candle flame, or an ordinary gas flame, will deposit a layer of carbon on any substance with which it comes in contact; and, as the flame contains solid particles of unburnt carbon, we may speak of the combustion as being incomplete. But, in the Bunsen burner, sufficient air is admitted to entirely burn up these carbon particles—that is, to convert them all into carbonic acid gas. Consequently, little light is emitted;

but as the combustion is complete, the heat of the flame is more intense. All our gas heating and cooking stoves are constructed on the principle of the Bunsen burner.

Experiment 74.—The Bunsen flame may be proved to be hollow by lowering a piece of wire gauze or a piece of stout white blotting-paper into the broadest portion of the flame. If blotting-paper be used it should be kept steady in the flame for only a moment, and, on removing it, a black ring will be seen, as in Experiment 72. The wire gauze may be kept in the flame for a longer time, when a white-hot ring will be produced, surrounding a dark centre. It will also be noticed that the gas burns below the gauze, but not above it; yet, if a light be applied above, the gas which passes through from the hollow part of the flame underneath will be ignited. The gas did not burn above the gauze in the first instance

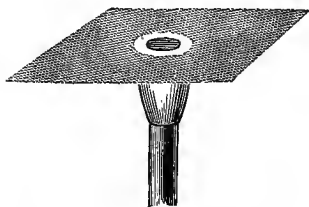


FIG. 42.—ILLUSTRATING THE HOLLOWNESS OF THE BUNSEN FLAME.

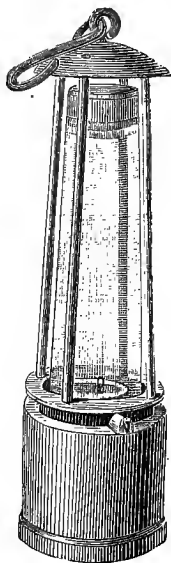


FIG. 43.—THE SAFETY LAMP

because the metal conducted away so much heat that the temperature did not rise high enough to ignite it. The **miner's safety lamp** is constructed on this principle. The flame is completely surrounded by a close mesh of wire gauze, so that any explosive gases which may exist in the mine, although they may burn *inside* the lamp after penetrating the gauze, yet they are not ignited outside.

112. The blowpipe flame.

Experiment 75.—Pass a current of air into a candle flame by means of a blowpipe. The luminosity of the flame will be



FIG. 44.—THE BLOWPIPE FLAME.

1, The flame, consisting of two distinct portions ; 2, nozzle of the blowpipe.

greatly weakened, as in the Bunsen burner, owing to the complete combustion of the carbon of the tallow or wax. The pale blue flame thus produced is hotter than the ordinary

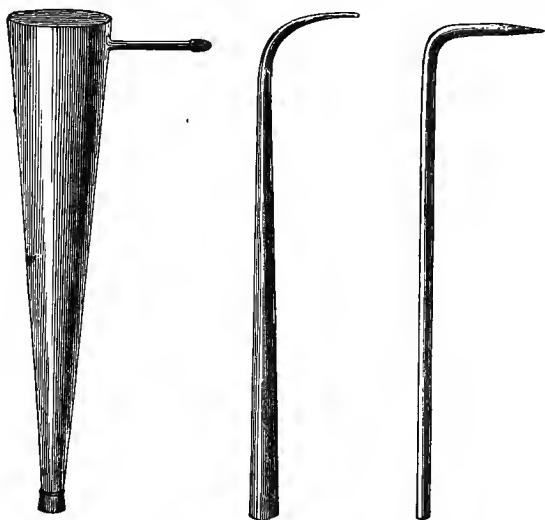


FIG. 45.—VARIOUS FORMS OF BLOWPIPES.

candle flame, and is commonly known as the blowpipe flame. The blowpipe may be used in this manner with the gas flame, or with the flame of any kind of lamp.

113. Summary.

1. When coal is distilled, coal-gas and various other substances are given off, and coke remains.
2. Coal-gas is similar to hydrogen in many of its properties, and is often used as a substitute for that gas.
3. Coal-gas contains hydrogen and carbon. When it burns these elements combine with the oxygen of the atmosphere, forming respectively water and carbonic acid gas.
4. The chief constituents of coal-gas are heavy carburetted hydrogen (olefiant gas) and light carburetted hydrogen (marsh gas or fire-damp).
5. Tar, pitch, naphtha, benzine, and carbolic acid are some of the products of the distillation of coal.
6. Our chief supplies of the ammonia salts are also derived from the ammoniacal liquor of the gas-works.
7. A flame consists of a hollow core of unburnt gas, surrounded by a more or less luminous portion.
8. The light of a bright flame is emitted from particles of white-hot, unburnt carbon.
9. The pale blue portion of a flame is produced by the combustion of hydrogen.
10. The flame of the Bunsen burner and the blowpipe flame are almost non-luminous because the combustion of the carbon is complete, all of it being converted into carbonic acid gas by combination with the abundant supply of oxygen.

QUESTIONS ON LESSON XII.

1. Describe an experiment which illustrates the manufacture of coal-gas.
2. What is *coke*? How is it obtained?
3. What are the chief properties of coal-gas? In what respects does it resemble hydrogen?
4. Describe a few experiments by which you would illustrate the properties of coal-gas.
5. What do you know of the chemical composition of coal-gas? How would you prove that it contains hydrogen and carbon?
6. Name some of the chief products of the distillation of coal.
7. Describe the structure of a candle flame.
8. How would you prove that the candle flame contains a core of unburnt gas?
9. What is the cause of the light emitted from a candle flame?

10. Describe the construction and principle of the Bunsen burner.
 11. In what respects does the combustion of an ordinary gas flame differ from that of the Bunsen flame?
 12. What is the blowpipe flame? What are its special advantages?

LESSON XIII.

SULPHUR.

114. Occurrence in nature.

Sulphur or **Brimstone** is an elementary substance found largely in the earth's crust. In volcanic regions it occurs in a free state, that is, not in combination with other elements. This *native* sulphur consists of masses of yellow, transparent crystals. Sulphur is also found largely in combination with various metals as **sulphides**, and in combination with metals and oxygen as **sulphates**. The following are a few of these **native sulphur compounds** :—

Iron pyrites . . .	Iron sulphide . . .	} Each containing the metal or metals with sulphur.
Copper pyrites . . .	Iron and copper sulphide . . .	
Blende . . .	Zinc sulphide . . .	
Galena . . .	Lead sulphide . . .	
Cinnabar . . .	Mercury sulphide . . .	} Each containing the metal with sulphur and oxygen.
Gypsum . . .	Calcium sulphate . . .	
Heavy spar . . .	Barium sulphate . . .	
Epsom salts . . .	Magnesium sulphate . . .	

Sulphur also exists in smaller quantities in parts of animals and plants.

115. Pure sulphur may exist in several different forms.

In addition to the native crystals above mentioned, the following varieties are interesting and important :—

(1) **Roll sulphur**, prepared by melting sulphur and casting it in moulds.

(2) **Flowers of sulphur**, consisting of a very fine crystalline powder.

(3) **Prismatic, needle-like crystals**.

(4) **Plastic sulphur.**

We shall presently learn how the last three varieties may be prepared.

116. The effects of heat on sulphur.

Experiment 76.—Place a little sulphur on a piece of tin plate, and then apply heat by supporting it over the flame of a spirit lamp or gas burner. The sulphur soon melts, becomes darker in colour, and finally catches fire, burning with a pale blue flame, and giving off a gas with a suffocating odour. Now heat a little sulphur in a test-tube, and note more particularly the changes which take place. First the sulphur melts (at 115° C.), producing an amber-coloured liquid. After a short time it thickens (at about 230° C.) and becomes darker in colour. It continues in this state until the temperature rises to about 250° , when it becomes thin again, and remains as a reddish-black liquid till (at 440°) it boils, giving off a dark red vapour. It will now be noticed that, when this red vapour comes in contact with the cool air, it is converted back into the solid state in the form of a fine powder, some of it settling on the sides of the glass. This deposit consists of **flowers of sulphur**. Flowers of sulphur are prepared, on a large scale, by boiling crude sulphur in large retorts, and condensing the vapour quickly by passing it into cold chambers.

117. To prepare needle-like crystals of sulphur.

Experiment 77.—Melt about half a pound of sulphur (roll sulphur is the best) slowly in a clay crucible or earthenware pot, being careful not to let it boil. When all is melted, set it aside to cool. As soon as a firm crust has formed on the top, make a hole in it with a penknife, and pour out the liquid interior. Now cut out the top crust, and a mass of beautiful, transparent needles will be seen. These crystals soon become yellow and opaque.

118. To prepare plastic sulphur.

Experiment 78.—Heat about an ounce of sulphur in a clay crucible till it has all melted and been converted into a thick syrupy liquid. Now pour it into a vessel of cold water.

It will then be found to be a soft plastic mass, easily drawn out into threads, and somewhat resembling caoutchouc ; but in the course of an hour or two it again becomes brittle.

119. The solution of sulphur, and the preparation of crystals from the solution.

Experiment 79.—Sulphur is insoluble in water, but is soluble in a liquid called carbon disulphide. Place a few small fragments of roll sulphur in a bottle containing some carbon disulphide, and shake it till all the sulphur has dissolved. Then pour some of this solution of sulphur into a watch-glass or shallow basin, and allow the liquid to evaporate without heat. As the liquid disappears, it deposits a mass of crystals of sulphur. It is interesting to note that plastic sulphur is insoluble in carbon disulphide.

120. Iron and copper sulphides.

Experiment 80.—Mix 7 drachms of fine iron filings with 4 drachms of flowers of sulphur, and throw the mixture, a little at a time, into a red-hot crucible. Under the influence of the heat, the two elementary substances combine, forming a black compound, called **iron sulphide**. This sulphide is similar in composition to the native iron pyrites, but does not contain the elements in the same proportion. If we mix the iron and sulphur in any other proportions, they still combine in the proportions 7 : 4, leaving the excess of one or the other in a free state between the particles of the sulphide, or giving it up to the oxygen of the air.

Experiment 81.—Crush some **iron pyrites** and heat it strongly in a test-tube, holding the tube in a horizontal position. After a short time sulphur will be driven off in the form of vapour, and this condenses, forming a yellowish deposit on the cool part of the tube.

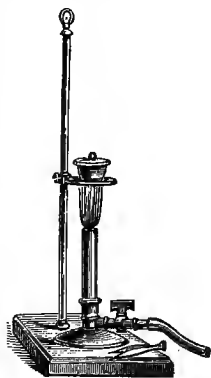


FIG. 46.—BUNSEN BURNER AND CRUCIBLE.

Experiment 82.—Boil a little sulphur in a test-tube, and pass into the tube a spiral of fine copper wire, or some copper turnings. As soon as the metal comes in contact with the sulphur vapour, it will burn with incandescence, and be converted into a black, brittle **copper sulphide**.

121. Summary.

1. The element sulphur is widely distributed in nature. It occurs in the free state, and as sulphides and sulphates.

2. There are several varieties of pure sulphur.

3. Sulphur melts at 115° C., thickens at 230° , becomes fluid again at 250° , boils at 440° , giving off a dark-red vapour.

4. Flowers of sulphur are prepared by boiling sulphur and rapidly cooling its vapour.

5. Prismatic needles of sulphur may be prepared by melting sulphur and allowing it to cool slowly.

6. Plastic sulphur is produced by pouring sulphur at 230° C. into cold water. Its plasticity is not permanent.

7. Sulphur is soluble in carbon disulphide; and the solution yields crystals on evaporating.

8. Iron and sulphur, mixed and heated, produce iron sulphide.

9. Iron pyrites yields sulphur when heated.

10. Copper burns in sulphur vapour, forming copper sulphide.

QUESTIONS ON LESSON XIII.

1. In what different forms does sulphur exist in nature? Give the names of some of the more common native compounds containing sulphur.

2. Describe the different varieties of sulphur.

3. Describe the changes which take place when sulphur is heated to its boiling point.

4. How are flowers of sulphur prepared?

5. How would you prepare prismatic crystals of sulphur?

6. Describe plastic sulphur. How would you prepare it?

7. How would you dissolve sulphur?

8. Describe a method of preparing sulphur crystals without heat.

9. How would you prepare iron sulphide?

10. Give an experimental proof that iron pyrites contains sulphur.

11. Describe a method of preparing copper sulphide.

12. Write a summary of the chief properties of sulphur.

LESSON XIV.

SULPHUR DIOXIDE AND SULPHUROUS ACID.

122. Combustion of sulphur.

WHEN sulphur is burnt in air or oxygen, it combines with its own weight of oxygen, and thus forms a compound gas called **sulphur dioxide** or **sulphurous acid gas**. The suffocating odour of burning sulphur is due to this gas.

123. Preparation of sulphur dioxide.

Experiment 83.—This gas may be conveniently prepared by heating sulphuric acid with fragments of copper in a glass flask provided with a delivery tube. Some of the sulphuric acid combines with the copper, forming copper sulphate, while the remainder gives up sulphur dioxide. It is much heavier than air (density = 32), and may therefore be collected by the downward displacement of air. A few jars should be filled with it and set aside for the following experiments.

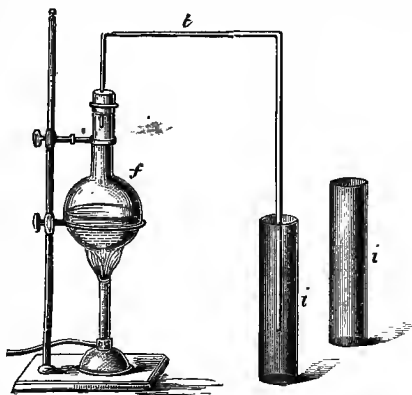


FIG. 47.—APPARATUS FOR PREPARING SULPHUR DIOXIDE.

f, flask containing copper and sulphuric acid ;
d, delivery tube ; *i*, tall gas jars.

When the gas ceases to come off, filter the liquid remaining in the flask, then transfer it to a shallow vessel and evaporate it at a low temperature. As the evaporation proceeds, blue crystals of copper sulphate will be deposited.

124. Sulphur dioxide dissolves in water.

Experiment 84.—Invert one of the jars of the gas in a vessel of water. The water rapidly dissolves the gas, and then rises in the jar to take its place. If the gas is pure and free from air, the water will rise till it quite fills the jar. Now add a little solution of litmus to the water in which the sulphur dioxide is dissolved; it immediately turns red, proving the presence of an acid. This acid is called **sulphurous acid**, and is composed of sulphur dioxide and water. A stronger acid may be produced by shaking up a smaller quantity of water with the gas in a jar. Water will dissolve about 50 times its own volume of sulphur dioxide when cold, but the gas is again liberated when heat is applied to the solution.

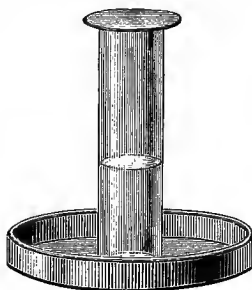


FIG. 48.—ILLUSTRATING THE SOLUBILITY OF SULPHUR DIOXIDE.

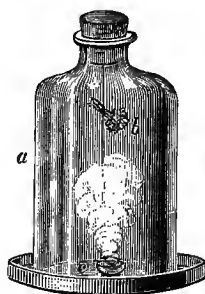


FIG. 49.—ILLUSTRATING THE BLEACHING POWER OF SULPHUR DIOXIDE.
a, gas jar; *b*, bunch of violets; *c*, crucible containing burning sulphur.

125. Bleaching power of sulphur dioxide.

Experiment 85.—Suspend a bunch of violets or a red rose in a jar of the gas. The colour will soon be destroyed. When a red rose is bleached in this manner its colour may be restored by dipping it into dilute sulphuric acid. This experiment may be varied by burning a little sulphur in a small saucer or crucible underneath a bell-jar in which the flowers are suspended, taking care that the flame is not so near to the flowers as to scorch

them. Many vegetable colours may be bleached by means of this gas.

126. Sulphur dioxide is not a supporter of combustion. A lighted taper placed in the gas is immediately extinguished. It is not combustible.

127. Summary.

1. Sulphur dioxide is a compound gas, composed of sulphur and oxygen.

2. It may be prepared by burning sulphur in air or oxygen, or by heating sulphuric acid with copper.

3. It is soluble in water, and forms with the water sulphurous acid.

4. Sulphur dioxide has the power of bleaching vegetable colouring substances.

5. It is not a supporter of combustion, nor is it combustible.

QUESTIONS ON LESSON XIV.

1. What chemical change takes place when sulphur is burnt?
2. Describe a method of preparing and collecting sulphur dioxide. Sketch the apparatus you would use.
3. How would you prove that this gas is soluble in water?
4. Describe the properties of the substance formed when sulphur dioxide is dissolved in water.
5. How would you illustrate the bleaching power of sulphur dioxide?
6. Enumerate the chief properties of sulphur dioxide.

LESSON XV.

SULPHUR TRIOXIDE AND SULPHURIC ACID.

128. Sulphur trioxide.

Experiment 86.—If sulphur dioxide be mixed with oxygen, the two gases will not combine; but if the mixture be passed over heated platinised asbestos, they will, forming white solid fumes of sulphur trioxide. This substance contains 40 parts by weight of sulphur to 60 of oxygen. Sulphur trioxide has a

great affinity for water, with which it combines with a hissing noise, forming sulphuric acid.

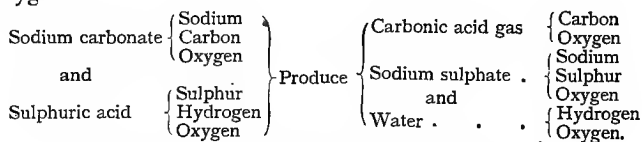
129. Sulphuric acid.

Pure sulphuric acid is a colourless, thick, oily liquid, having a specific gravity of 1.8. Under the name of **oil of vitriol**, it is largely used in various manufactures and industries. One pint of water weighs $1\frac{1}{4}$ lb., therefore a corresponding volume of sulphuric acid weighs $1\frac{1}{4}$ lb. \times 1.8 = $2\frac{1}{4}$ lbs. It has a powerful acid taste, and readily turns blue litmus red, even when diluted with 20 or 30 times its volume of water.

130. Action of sulphuric acid on carbonates.

Like all other acids, this one decomposes carbonates, setting free carbonic acid gas, and forming new compounds called sulphates. The following experiment is an illustration of this:—

Experiment 87.—Put a little sodium carbonate (carbonate of soda) into a glass vessel, and add some diluted sulphuric acid, a little at a time, till all the carbonate has disappeared, and no further effervescence takes place. The gas which passes off will not support the combustion of a taper, and if it be passed into lime-water, produces a precipitate of calcium carbonate (see Art. 103), thus proving it to be carbonic acid gas. Now allow the liquid to evaporate slowly in a shallow vessel, and when all has disappeared a crystalline substance remains, which is sodium sulphate or sulphate of soda, composed of the metal sodium in combination with sulphur and oxygen:—



131. Sulphuric acid has great affinity for water.

Experiment 88.—Pour four ounces of sulphuric acid into a glass vessel and add gradually one ounce of water, stirring all the time with a glass rod. Great heat is evolved, thus illus-

trating the powerful chemical affinity existing between the two substances. Now stand in the mixture a test-tube containing alcohol or methylated spirit, and after a few minutes the spirit will boil vigorously.

132. Sulphuric acid destroys organic matter.

Experiment 89.—The following experiment also illustrates the affinity of sulphuric acid for the elements of water:—Moisten about an ounce of loaf sugar with boiling water, and then add quickly about an ounce or more of strong sulphuric acid. A black mass of carbon is almost immediately set free from the sugar. Sugar contains the three elements carbon, hydrogen, and oxygen; the acid combines with the elements of water (hydrogen and oxygen), and sets the carbon free. A similar action takes place when wood or cork comes in contact with this acid, or when oil and sulphuric acid are shaken together.

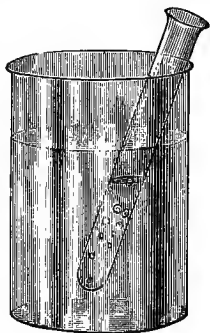


FIG. 50. — ILLUSTRATING THE AFFINITY OF SULPHURIC ACID FOR WATER.

133. Summary.

1. Sulphur trioxide is prepared by strongly heating a mixture of sulphur dioxide and oxygen.
2. It combines with water, forming sulphuric acid.
3. Sulphuric acid has a sharp sour taste, and reddens blue litmus. Its specific gravity is 1.8.
4. Sulphuric acid decomposes carbonates, liberating carbonic acid gas and forming sulphates.
5. It has a powerful affinity for the elements of water, and destroys organic matter by depriving it of these elements and setting free its carbon.

QUESTIONS ON LESSON XV.

1. How does sulphur dioxide differ in composition and properties from sulphur trioxide?
2. Describe a method of preparing the latter.
3. How may sulphur trioxide be converted into sulphuric acid?
4. What is the weight of (1) a gallon of water, and (2) a gallon of sulphuric acid?

5. In what respects do sulphurous and sulphuric acids resemble each other ; and how do they differ ?
 6. Describe fully the action of sulphuric acid on carbonate of soda. What new compounds are formed ? Describe these compounds.
 7. What are the chemical elements in
 - (a) Sulphuric acid
 - (b) Sodium carbonate
 - (c) Sodium sulphate ?
 8. Describe an experiment which illustrates the affinity of sulphuric acid for water.
 9. Describe fully the action which takes place when sulphuric acid is brought in contact with organic matter. Illustrate by an experiment.
-

LESSON XVI.

COMMON SALT AND CHLORINE.

134. Common salt.

COMMON table salt is a compound substance containing the elements sodium and chlorine. Sodium is a metal to be hereafter described, and chlorine is a yellowish-green gas (Gr. *chloros*, green). A compound composed of an element with chlorine is called a chloride. Common salt is called chloride of sodium, or sodium chloride. It is found largely in nature. It occurs as extensive beds of a colourless, red, blue, or purple substance known as **rock-salt**. It is also obtained largely from sea-water by evaporation. The salts of other metals beside sodium are found in rock-salt and sea-salt in small proportions.

135. Preparation of chlorine from common salt.

Experiment 90.—Fit a large glass flask with a cork and bent delivery tube (see fig. 51). Mix thoroughly 2 ozs. of common salt and 2 ozs. of the black oxide of manganese. Introduce this mixture into the flask, and add sufficient sulphuric acid (which has been previously diluted with its own volume of water) to make a thin paste. Shake these substances well together till they are thoroughly mixed, and no dry place appears at the bottom of the flask. Now apply a gentle heat, and chlorine gas will soon come off with great rapidity, and may be recognised by its peculiar odour. Care must be taken to avoid inhaling this gas, as it is a violent irritant. If the experiments are performed in a room, see that it is well ventilated,

and allow none of the gas to escape unnecessarily. Chlorine gas is about $2\frac{1}{2}$ times heavier than air, and $35\frac{1}{2}$ times heavier than hydrogen. It may, therefore, be collected by the downward displacement of air. Again, as the gas is of a yellowish-green colour, one may see when the jars are full, if prepared by daylight. But if prepared by artificial light, it will be advisable to collect the gas over water in the pneumatic trough; but warm water must be used, as chlorine is soluble in cold water. As soon as each jar is full, replace it quickly by another, and cover it at once to prevent the escape of gas. About six jars of chlorine will be necessary for the illustration of its more important properties.

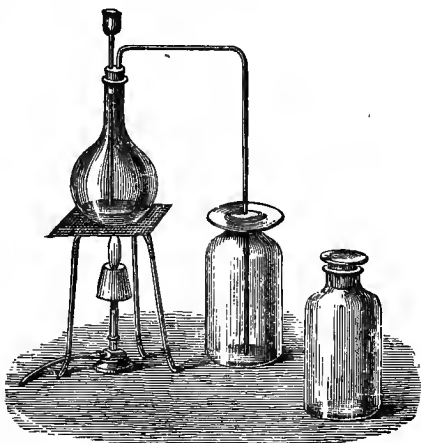


FIG. 51.—APPARATUS FOR PREPARING AND COLLECTING CHLORINE.

136. Preparation of chlorine from hydrochloric acid.

Experiment 91.—Hydrochloric acid¹ is composed of the elements hydrogen and chlorine. Mix about two ounces of the oxide of manganese with sufficient hydrochloric acid to form a thin paste, and heat the mixture gently, using the same arrangement of apparatus recommended for the last experiment. The chlorine produced by this method is liberated from the hydrochloric acid.

137. Spontaneous combustion of phosphorus in chlorine.

Experiment 92.—Take a piece of phosphorus about as large as a pea, dry it with blotting-paper, place it in a defla-

¹ Hydrochloric acid is really a gas. The liquid known as hydrochloric acid, or spirits of salt, is a solution of this gas in water.

grating spoon, and introduce it into one of the jars of chlorine. The phosphorus ignites, and burns with a feeble, greenish flame, combining with the chlorine, and forming a yellow compound called chloride of phosphorus. As soon as the combustion is over, remove the jar into the open air, and leave it open to allow of the escape of the unused gas. This remark applies also to all the other experiments with chlorine.

138. Spontaneous combustion of copper in chlorine.

Experiment 93.—Pour into another bottle of chlorine just sufficient water to cover the bottom. Drop into the jar some very fine brass¹ or copper filings. As the particles pass through the gas they burn brilliantly, the copper combining with the chlorine and forming green copper chloride. Instead of the filings copper leaf or dutch-metal leaf (which contains much copper) may be used. The water is placed in the jar to prevent fracture, which might be caused by the particles of burning metal coming in contact with the glass.

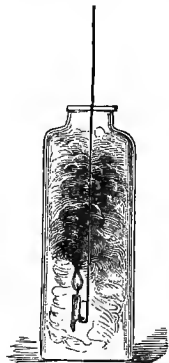


FIG. 52.—A CANDLE
BURNING IN A JAR
OF CHLORINE.

139. Antimony burns spontaneously in chlorine.

Experiment 94.—Antimony is a crystalline, brittle metal which may be powdered in a mortar. Reduce a little of this metal to a fine powder, and let it fall into a jar of chlorine, in which has been placed previously a little water as in the last experiment. The particles of antimony burn brilliantly, producing the chloride of the metal.

140. Combustion of a candle in chlorine.

Experiment 95.—Place a lighted candle or taper in a jar of chlorine. The flame is not extinguished, but the candle or taper burns with a very feeble flame, giving off much carbon in the form of smoke. The 'wax' or tallow of the candle contains the elements carbon and hydrogen. Chlorine supports the

¹ Brass is a mixture of copper and zinc.

combustion of hydrogen, but not of carbon. As the candle burns, its hydrogen combines with the chlorine, forming hydrochloric acid ; but the carbon, having little chemical affinity for chlorine, is set free. As this experiment proceeds, it will be seen that the yellowish-green gas (chlorine) gradually disappears, giving place to a colourless gas (hydrochloric acid) which forms white fumes on exposure to the air.

141. Combustion of turpentine in chlorine.

Experiment 96.—Warm a little oil of turpentine in a test-tube, being careful not to allow the flame of the spirit lamp or gas burner to get near the mouth of the tube. Saturate a strip of blotting-paper with this warm turpentine and drop it into a jar of chlorine. The turpentine immediately bursts into a flame and gives off dense volumes of smoke. Oil of turpentine contains carbon and hydrogen. The hydrogen and chlorine combine, forming hydrochloric acid, and the carbon is set free. Thus the chemical action is exactly the same as that exhibited in the last experiment.

142. Combustion of hydrogen in chlorine.

Experiment 97.—Pass a jet of burning hydrogen into a jar of chlorine. The two gases combine, forming hydrochloric acid ; and as soon as the chlorine has been converted into hydrochloric acid, the flame is extinguished, for hydrochloric acid will not support the combustion of hydrogen.



FIG. 53.—ILLUSTRATING THE COMBUSTION OF HYDROGEN IN CHLORINE.

143. Formation of common salt.

Experiment 98.—The metal sodium burns in chlorine when strongly heated, and the formation of common salt is the result

of the combination of the two elements. Common salt is composed of 23 parts *by weight* of sodium and $35\frac{1}{2}$ parts of chlorine.

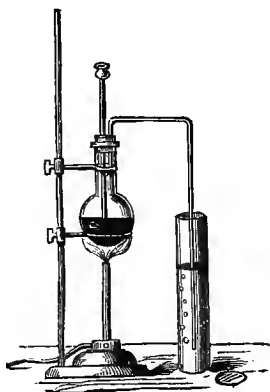


FIG. 54.—SHOWING THE METHOD OF PREPARING A SOLUTION OF CHLORINE.

144. Solution of chlorine.

Experiment 99.—Water will dissolve more than twice its volume of chlorine. A solution of chlorine may be prepared by shaking the water and the gas together in a jar or bottle ; or it may be prepared by passing the gas direct from the generating apparatus into water. The solution has the same colour and odour as the gas, and possesses many of the other properties of chlorine.

145. Bleaching power of chlorine.

Chlorine gas is a powerful bleaching agent. It has the property of destroying organic colouring matter in the presence of water. Dry chlorine will not bleach ; and on this account a solution of chlorine should be used, or the substance to be bleached may be first moistened, and then suspended in a jar of chlorine gas.

Experiment 100.—This property of chlorine may be conveniently illustrated by the following substances :—

Indigo	Writing ink
Litmus	Violets
Turkey red	Red rose

or any vegetable colours. Colour once destroyed by chlorine can never be restored.

146. Bleaching powder.

Bleaching powder, or **chloride of lime**, is a combination of lime and chlorine, and therefore contains calcium, oxygen, and chlorine. This compound readily parts with its chlorine, and is, on this account, largely used as a bleaching agent as well as a disinfectant. The following experiment illustrates the method of **bleaching on a large scale** :—

Experiment 101.—Prepare two baths, one containing a solution of bleaching powder, and the other very dilute sulphuric acid. Rinse the article to be bleached first in the solution of bleaching powder, and then in the acid bath. Repeat this process, using the baths alternately, till the colour is entirely destroyed. The acid is used to set free the chlorine in the bleaching powder.

147. Summary.

1. Common salt is composed of sodium and chlorine. Rock salt and sea salt are impure varieties of common salt.
2. Chlorine may be prepared by heating a mixture of
 - (a) common salt, oxide of manganese and sulphuric acid,
 - (b) oxide of manganese and hydrochloric acid.
3. Phosphorus, copper, and antimony burn spontaneously in chlorine, forming chlorides.
4. A candle and turpentine will burn in chlorine (the latter spontaneously), forming hydrochloric acid and liberating carbon.
5. Chlorine supports the combustion of hydrogen, forming hydrochloric acid.
6. Sodium, when heated, burns in chlorine forming common salt.
7. Chlorine is soluble in water. The solution possesses many of the properties of the gas.
8. Chlorine has the power of destroying organic colours.
9. Bleaching powder is a combination of lime with chlorine. The chlorine is set free from it by an acid.

QUESTIONS ON LESSON XVI.

1. What is common salt? How does it occur in nature?
2. Describe fully a method of preparing chlorine from common salt. How would you collect the gas?
3. How would you obtain chlorine from hydrochloric acid?
4. Enumerate the chief properties of chlorine.
5. Give illustrations of the spontaneous combustion of substances in chlorine.
6. What is copper chloride, and how would you make a small quantity of it?
7. Describe fully the chemical change which takes place when a lighted taper is placed in chlorine gas.

8. Describe the action of chlorine on oil of turpentine.
9. How would you form hydrochloric acid by the direct union of its two elements?
10. How would you prepare common salt by the direct union of its elements?
11. How would you prove that chlorine is soluble in water?
12. In what respects do chlorine and sulphur dioxide resemble each other? In what respects do they differ?
13. Describe an experiment which proves that chlorine has but little affinity for carbon.
14. How would you bleach a bunch of violets by means of a jar of chlorine gas?
15. Name several substances which may be bleached by chlorine.
16. What is bleaching powder?
17. Describe the method of bleaching with chloride of lime on a large scale.
18. You are provided with four jars of chlorine. Say what you would do with them in order to illustrate the characteristic properties of the gas.

LESSON XVII.

ACIDS, BASES, SALTS, AND ALKALIES.

148. General properties of acids.

SOME of the characteristics of acids have already been noticed. These will be repeated here, together with others, in order to provide the student with a complete definition of the term acid:—

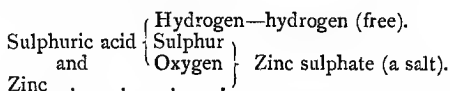
(a) **Acids possess a sharp sour taste.**

(b) **They all have the power of turning blue litmus red.** Other vegetable blues are also turned red by acids. Thus, vinegar turns purple cabbage to a bright red; and blue flowers may be similarly changed in colour by immersion in an acid.

(c) **All acids have the power of decomposing carbonates,** and setting free the carbonic acid gas they contain. (See Art. 101.)

(d) **All acids contain hydrogen,** which is readily displaced by metals. An example of this property has already been noticed in Art. 82. In this case, sulphuric acid is added to the metal zinc; the hydrogen of the acid is liberated; and

the zinc, taking the place of this hydrogen, forms zinc sulphate with the sulphur and oxygen :—

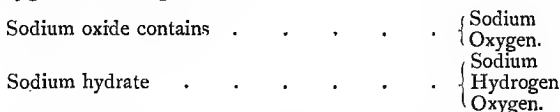


When an acid gives up its hydrogen for a metal, a substance called a salt is formed. Further information concerning salts will be given below (Art. 150).

(e) **All acids have the power of neutralising alkalies.**
(See Art. 152.)

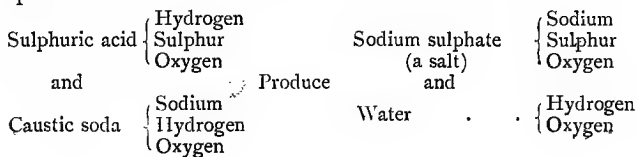
149. Bases.

A base is a compound which, when acted on by an acid, forms a salt. Among the most important bases are the oxides and hydrates of many metals. The **oxide** of a metal is the compound produced by the chemical union of the metal with oxygen gas. The **hydrate** of a metal is produced by its combination with the elements of water. A hydrate of a metal consists, therefore, of three elements, viz. the metal, hydrogen, and oxygen. Examples :—



Experiment 102.—The following experiment illustrates the formation of a salt by the action of an acid on a base :—

Make a solution of sodium hydrate (caustic soda) in water by dissolving about one drachm of the hydrate in $\frac{1}{2}$ oz. of water. Then add sulphuric acid, a little at a time, till the liquid ceases to turn red litmus blue. Now evaporate slowly in a shallow vessel till crystals are deposited. These are crystals of sodium sulphate :—



150. Salts.

A salt is a compound formed by the action of an acid on a base. The formation of a salt in this way has already been illustrated in our last experiment. Numerous other illustrations may be chosen by selecting various other acids and bases. Every acid has its corresponding salts. Thus :—

Acids		Salts
Sulphuric acid, acting on bases	forms	sulphates.
Nitric acid,	„	nitrates.
Hydrochloric acid,	„	chlorides.
Carbonic acid,	„	carbonates.
Phosphoric acid,	„	phosphates.
Hydrosulphuric acid,	„	sulphides.
Acetic acid,	„	acetates.
Tartaric acid,	„	tartrates, &c.

It will be noticed that, when the acid contains oxygen, the termination *ate* is applied to the name of its series of salts, which also contain oxygen; but when the acid contains no oxygen, neither do its salts, and in this case the termination *ide* is used :—

Nitric acid	{ Hydrogen Nitrogen Oxygen }	form a nitrate	{ Metal Nitrogen Oxygen }
and			
Metal			
Hydrochloric acid	{ Hydrogen Chlorine }	form a chloride	{ Metal Chlorine }
and			
Metal			

151. Alkalies.

Alkalies are compounds that have the power of turning red litmus blue, and of neutralising acids. They have a soapy taste, and absorb carbonic acid gas. The chief alkalies are :—

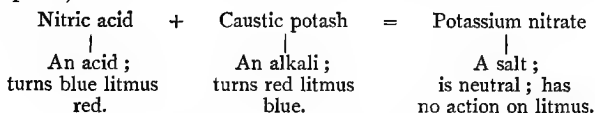
1. Sodium hydrate (caustic soda) . . . { Sodium
Hydrogen
Oxygen
2. Potassium hydrate (caustic potash) . . { Potassium
Hydrogen
Oxygen
3. Ammonium hydrate (liquid ammonia) . { Ammonium
Hydrogen
Oxygen.

Sodium and potassium are elementary metals; ammonium is a compound containing the elements nitrogen and oxygen.

152. Acid, alkaline, and neutral substances.

When a substance has the power of turning *blue litmus red* it is said to be **acid**, or to have an acid reaction. When it has the power to turn *red litmus blue*, it is said to be **alkaline**, or to have an alkaline reaction. If it has *no action on either red or blue litmus*, it is said to be **neutral**; that is, it does not partake of the properties of either acids or alkalies; this is illustrated by the following experiment:—

Experiment 103.—Dissolve a few grains of caustic potash in water in a test-tube: add to this a few drops of tincture of blue litmus. Of course the colour is not changed by the alkaline fluid. Into another test-tube place a small quantity of nitric acid (aquafortis), and dilute with water. Now add the diluted nitric acid to the alkaline liquid very slowly and carefully till the blue solution *just begins to turn red*. The fluid is now *neutral*, for it will not act decidedly on either red or blue litmus. Evaporate this liquid to dryness slowly, and crystals of *a salt* will remain. This salt is potassium nitrate (nitre, or saltpetre).—



153. The three chief acids.

(a) **Sulphuric acid**, which has already been described in Arts. 129–133.

(b) **Nitric acid**, known commercially as **aquafortis**, is a colourless liquid which gives off fumes when exposed to air. It is heavier than water, but not so heavy as sulphuric acid, its specific gravity being 1.5. It acts powerfully on organic matter, corroding it and staining it yellow. Nitric acid bleaches indigo, and dissolves most metals, forming nitrates.

Experiment 104.—Add nitric acid to copper, a greenish-blue liquid is produced and a reddish-brown gas is given off. The greenish-blue fluid is a solution of copper nitrate, from which the crystals of the salt may be obtained by evaporation.

(c) **Hydrochloric acid** (muriatic acid, spirit of salt) is really

a gas which may be prepared by heating a mixture of common salt and sulphuric acid. The hydrochloric acid of the shops is a strong solution of this gas in water, having a specific gravity of about 1.2. The solution possesses all the properties of the gas, which may be driven off from the liquid by the application of heat. This acid also fumes when exposed to air, and when pure is colourless; but, unlike nitric acid, it has very little action on copper. It dissolves many of the other metals, forming salts called **chlorides**.

154. Summary.

1. Acids :—

- (a) Have a sharp taste.
- (b) Turn vegetable blues red.
- (c) Decompose carbonates, setting free carbonic acid gas.
- (d) Contain hydrogen.
- (e) Form salts with bases.
- (f) Neutralise alkalies.

2. A **base** is a compound (generally an oxide or a hydrate) which forms a salt with an acid.

3. A **salt** is the product of an acid and a base.

4. When an acid gives up its hydrogen for a metal, a salt is formed.

5. **Alkalies** have contrary properties to acids; and they neutralise acids, forming salts.

6. The following table gives a summary of the **distinguishing properties of the three chief acids** :—

Sulphuric acid (Oil of vitriol)	Nitric acid (Aquafortis)	Hydrochloric acid (Muriatic acid)
Colourless liquid. S.G. = 1.8 Does not fume in air. No action on indigo.	Colourless liquid. S.G. = 1.5 Fumes in air. Bleaches indigo.	Colourless gas, soluble in water. Fumes in air. No action on indigo.
Dissolves copper when heated, giving off sulphur dioxide, and forming copper sulphate. Chars organic matter.	Dissolves copper without heat, giving off reddish-brown fumes, and forming copper nitrate. Stains organic matter yellow.	No action on copper. Gives off chlorine when heated with manganese oxide.

QUESTIONS ON LESSON XVII.

1. What are the properties common to all acids?
 2. What changes take place when (1) nitric acid and (2) hydrochloric acid are added to sodium carbonate?
 3. What is a base? Give examples of bases.
 4. Describe fully the action of sulphuric acid on caustic soda, giving the names of the elements contained in the new compounds formed.
 5. What is a salt? Name some examples of salts.
 6. How would you prepare potassium sulphate, sodium nitrate, and sodium chloride?
 7. What is an alkali? Give examples of alkalies.
 8. Describe fully what is meant by the terms *acid*, *alkaline*, and *neutral*.
 9. How would you prepare a neutral solution by means of an acid and an alkali?
 10. Two bottles of acid are given you; one bottle contains sulphuric acid, and the other nitric. How will you distinguish between them?
 11. How would you distinguish between hydrochloric and nitric acids?
 12. What are the distinguishing characteristics of sulphuric, nitric, and hydrochloric acids? Write your answer in the form of a table.
-

LESSON XVIII.

AMMONIA.

155. Preparation of ammonia.

Experiment 105.—Mix well together about one ounce of powdered and dried sal-ammoniac (ammonium chloride) and an ounce of powdered quicklime. Transfer the mixture quickly to an 8-ounce flask or retort, and cover it with a layer of powdered quicklime only, to a depth of one inch or more. Fit the flask or retort with a cork and straight delivery tube, and apply a very gentle heat. The ammonia which passes off is a gas with a very powerful pungent odour. It is lighter than air, and may therefore be collected in inverted jars by the upward displacement of the air. In order to ascertain when each jar is full, place a strip of yellow turmeric paper outside the jar, just above its mouth. If it is full, the gas escaping will rise around the mouth of the jar and quickly turn the test-paper brown. Fill about three jars with the gas, close them, and set aside for the following experiments. The chemical action which

goes on in the flask is this :—ammonium contains nitrogen (82·3 parts) and hydrogen (17·7 parts), and this compound, with chlorine, forms sal-ammoniac. Lime contains calcium and oxygen. The calcium combines with the chlorine, forming calcium chloride, and the ammonia gas is set free from the sal-ammoniac. Water is also formed in the flask, and the use of

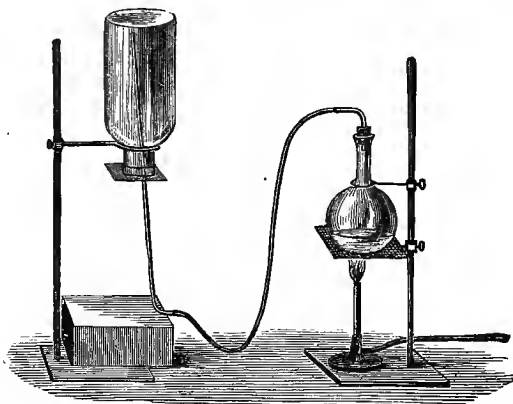


FIG. 55.—APPARATUS FOR THE PREPARATION AND COLLECTION OF AMMONIA GAS.

the layer of quicklime on the top of the mixture is to absorb this water, and so reduce the risk of breakage. Should the ammonia not be required in a dry state, it may be prepared by heating a paste made of the same two substances with water. Soda or potash may be used instead of lime in the preparation of this gas.

156. Ammonia burns very feebly.

Experiment 106.—Remove the stopper from one of the jars of ammonia, still keeping it inverted, and pass into it a lighted taper. The taper will be extinguished, but the gas will show a tendency to burn with a feeble, greenish flame at the mouth of the jar.

157. Ammonia is a powerful alkali.

Experiment 107.—Pass a strip of reddened litmus paper into a jar of ammonia, it is immediately turned blue, thus illus-

trating the alkaline character of the gas. This property may be further illustrated with the same jar of gas, by using various vegetable blues—blue flowers, for example.

158. Ammonia is very soluble in water.

Experiment 108.—Invert one of the jars of ammonia in a basin of water, and remove the stopper. The water will rapidly dissolve the gas, and rise in the jar to take its place. If the jar contains pure ammonia unmixed with air, the water will entirely fill the jar. Add a little reddened solution of litmus to the solution of ammonia thus produced, and again notice the alkaline reaction. A very strong solution of ammonia may be prepared by passing the gas direct from the generating apparatus into water which is kept cool. In this way the water may be made to dissolve nearly 1000 times its own volume of ammonia. As the gas dissolves the liquid increases in volume. The 'liquid ammonia' of the shops is a strong solution prepared in this way. If this solution be evaporated the whole of the water and ammonia pass off, leaving no residue.

159. Action of sulphuric acid on ammonia.

Experiment 109.—Add dilute sulphuric acid, drop by drop, to a solution of ammonia coloured blue with litmus, till the liquid *just begins to turn red*. The acid and the alkali have neutralised each other, and a salt, called **ammonium sulphate**, is the result of the combination. This salt may be obtained as colourless crystals by slow evaporation.

160. Action of nitric acid on ammonia.

In a similar manner ammonium nitrate may be formed by neutralising ammonia with nitric acid. The formation of this salt may also be illustrated as follows :—

Experiment 110.—Take two similar gas jars ; moisten the interior surface of one with strong aquafortis, and of the other with a strong solution of ammonia. Place the two jars mouth to mouth. The fumes of nitric acid will combine with the ammonia gas escaping from its solution and form solid white fumes of ammonium nitrate.

161. Formation of sal-ammoniac.

Experiment 111.—Repeat the last experiment, using hydrochloric acid instead of nitric. The vapours combine, forming dense solid fumes of ammonium chloride or sal-ammoniac. This salt may also be formed by adding hydrochloric acid to ammonia solution till the liquid is neutral; and the sal-ammoniac thus prepared may be crystallised by evaporation. It is a white solid, soluble in water, and having no smell of ammonia.

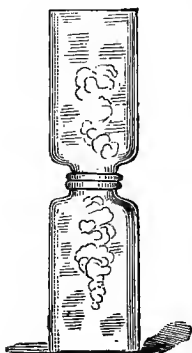


FIG. 56.—ILLUSTRATING
EXPERIMENTS 110 AND
111.

162 The salts of ammonium are volatile.

All the salts of ammonium are more or less volatile; that is, they have a tendency to pass away in the form of vapour. Some of them volatilise rapidly at ordinary temperatures; others only when moderately heated.

Experiment 112.—Heat a few crystals of ammonium nitrate in a test-tube. First the crystals melt; shortly afterwards the liquid boils, and passes off in the form of vapour. In this case the salt is decomposed into two simpler compounds—water and nitrous oxide (laughing gas)—the water passing off as steam, while the gas may be separated from this steam, and be collected, by passing it through cold water into an inverted jar filled with water.

Experiment 113.—A more remarkable result may be produced by heating a small quantity of sal-ammoniac in a test-tube. This salt volatilises without melting, and without undergoing any chemical change. It will all disappear from the bottom of the tube, and condense, like a deposit of hoar frost, on the upper, cool part. In this experiment the tube should be held in an oblique or nearly horizontal position, so that it may not become heated throughout its length.

163. Smelling salts and sal-volatile.

The substances sold under these names, and which smell so strongly of ammonia, contain carbonate of ammonium. They readily give off ammonia gas on exposure to air.

164. Ammonia produced by the decay of organic substances.

Ammonia is given off during the putrefaction of animal substances containing nitrogen. The gas is evolved slowly at ordinary temperatures, but the decomposition of the animal substance may be hastened by the application of heat, and still further assisted by adding caustic soda, potash, or lime.

Experiment 114.—As an illustration of this, crush a little cheese, add some caustic soda, and heat the mixture slowly. The escape of ammonia may soon be detected by its powerful odour. Ammonia is given off when the horns or hides of animals are heated, and it has been largely prepared from these substances; hence the name **spirits of hartshorn**. Sal-ammoniac is said to have been first prepared by the Arabs in the deserts of Libya, near the Temple of Jupiter Ammon, by heating waste animal matter. This explains the origin of the name ammonia. Coal contains nitrogen, and during the distillation of coal ammonia is given off with various other gases and vapours. Many of these vapours condense when cooled, forming a liquor containing ammonia. Sal-ammoniac is now mainly prepared by adding hydrochloric acid to this ammoniacal liquor of the gas-works.

Experiment 115.—By heating a little coal-dust in a test-tube, and passing the vapours evolved into a solution of red-denied litmus, we may prove the presence of an alkali. This alkali is ammonia.

165. Summary.

1. Ammonia gas may be prepared by heating a mixture of sal-ammoniac and quicklime.
2. Ammonia is very feebly combustible, and does not support combustion.
3. It is a powerful alkali, and a base.

4. It is very soluble in water. Its solution is called liquid ammonia, and ammonium hydrate.

5. Sulphuric, nitric, and hydrochloric acids all neutralise ammonia, forming respectively ammonium sulphate, nitrate, and chloride.

6. The salts of ammonia are readily decomposed.

7. Ammonium chloride volatilises when heated, without melting, and without undergoing chemical change.

8. Ammonia may be produced by the decomposition of animal substances containing nitrogen, such as horns, hides, hoofs, glue, isinglass, cheese, dried flesh, etc.

QUESTIONS ON LESSON XVIII.

1. How would you prepare and collect ammonia gas? Describe and sketch the apparatus you would use.

2. You are provided with four bottles of ammonia. Describe the experiments you would perform with them in order to illustrate the characteristic properties of the gas.

3. How would you prepare a solution of ammonia, and how would you obtain ammonia from it?

4. Describe experiments which illustrate the alkaline properties of ammonia.

5. What is the result of mixing sulphuric acid with a solution of ammonia?

6. How would you prepare ammonium nitrate?

7. How would you prepare sal-ammoniac?

8. Describe the changes which take place when the following substances are heated separately:—

(1) Ammonium nitrate.

(2) Sal-ammoniac.

(3) Ammonium hydrate.

9. What are 'smelling salts'?

10. Explain the origin of the terms:—

(1) Ammonia.

(2) Spirits of hartshorn.

11. From what source do we obtain our chief supplies of ammonia?

12. Describe an experiment by which ammonia may be obtained from an animal substance.

LESSON XIX.

CLAY.

166. Composition of clay.

CLAY is a rock substance composed of very fine particles, formed by the action of water and the atmosphere on certain hard rocks. These particles, after having been worn down, are carried by rivers or floods into the sea, or a lake, and there deposited as a sediment, sometimes accumulating till the bed thus formed is of considerable thickness. Clay deposited in this manner is often mixed with other substances carried along by the same stream ; but when pure, is composed entirely of two simpler compounds united chemically. These are **silica** and **alumina**. Clay is sometimes called **silicate of alumina**—a name which in itself reveals the chemical constitution of the substance. **Silica** is a compound containing the two elements **silicon** and **oxygen**. **Alumina** is also a binary compound, and is made up of the metal **aluminium** and **oxygen**. Thus there are three elements in clay—silicon, aluminium, and oxygen.

Pure clay	{ Silica	{ Silicon
						{ Alumina	{ Oxygen
							{ Aluminium
							{ Oxygen

The rocks known as slate and shale are formed from clay, and have the same chemical composition.

167. The properties of clay.

There are numerous varieties of clay, more or less pure, and varying in colour from a pure white to all the shades of yellow, red, blue, and purple. Among the purest varieties may be mentioned **kaolin** or **china-clay**, **pipe-clay**, and **fuller's earth**. Clay is capable of holding water for a long time, and may easily be moulded into various shapes. It is also hardened by baking, and is, therefore, well adapted for the manufacture of pottery and bricks. Vessels made of clay are usually glazed by fusing other substances on their surfaces ; this enables them

to hold liquids without allowing any to ooze through. Even the clay in its soft state may be made to hold water by well kneading it.

168. Silica.

The composition of silica has already been described. This compound is found largely in the earth's crust in a free state, and is known under various names. It occurs in granite rocks as irregular, colourless, glassy crystals. Sometimes it is found in the form of beautiful hexagonal crystals, and is then known as **rock crystal**, or quartz. **Flint** is also a variety of silica, as also are many of the stones used for ornamental purposes, such as **jasper**, **carnelian**, **amethyst**, **bloodstone**, **agate**, **onyx**, and **opal**.

169. Alumina.

Alumina is the oxide or rust of the metal aluminium. It does not exist largely in a free state, but chiefly in combination with silica and other substances. **Emery**, **ruby**, **oriental amethyst**, **sapphire**, **topaz**, and **emerald**, all contain a large percentage of alumina. Alumina may be prepared by heating ammonium alum.

170. Aluminium.

The metal aluminium is not found in a free state in nature. It is exceedingly light, having a specific gravity of 2.6. It is of a silvery white colour, takes a good polish, does not tarnish in air, is malleable and ductile, and is employed for the manufacture of useful and ornamental articles, when lightness, durability, and brilliancy are required. At present the metal is not manufactured in very large quantities, for, although so abundant, yet there is no known method of abstracting the metal from its compounds which is not very expensive.

171. Compounds containing alumina.

Among the most important of these compounds may be mentioned the **alums**. The alums are sulphates containing aluminium with another metal. The following is a list of the most important alums and their compositions:—

Potash alum	{ common alum }	.	Sulphate of aluminium and potassium.
Soda alum	.	.	„ „ sodium.
Ammonia alum	.	.	„ „ ammonium
Iron alum	.	.	„ „ iron.
Chrome alum	.	.	„ „ chromium.

Glass is a substance formed by fusing a mixture of silica lime, and soda ; or silica, oxide of lead, and potash. If the first mixture is used, plate-glass is produced ; the second mixture produces a variety known as flint glass. In glass-making the silica is generally used in the form of a fine white sand, such as we meet with in the beds of many of our rivers, and on the sea shore.

172. Summary.

1. Clay is composed of silica and alumina. Silica contains the elements silicon and oxygen : alumina contains aluminium and oxygen.

2. Aluminium is a light metal, resembling silver in appearance. It is malleable and ductile, and does not oxidise in air.

3. The various alums contain alumina. Common alum is a sulphate of alumina and potash.

4. Plate-glass is composed of silica, lime, and soda. Flint glass contains silica, oxide of lead, and potash.

QUESTIONS ON LESSON XIX.

1. What is clay ? How is it formed ?
 2. Describe the chemical composition of clay. Of what binary compounds is it composed ? What are the elements present in these binary compounds ?
 3. Describe the properties and uses of clay.
 4. What is silica ? Name some of its varieties.
 5. Describe the metal aluminium. What is the name given to its oxide ?
 6. What is the chemical composition of alum and of glass ?
-

LESSON XX.

THE METALS.

173. Elements.

ABOUT seventy elementary substances are known to chemists. Most of these are solid at ordinary temperatures, two are liquid, and four are gases. Some of these elements are found in nature in a free state ; but generally they are in combination with other elements. There are resemblances among them, but each one has properties peculiar to itself, by which it may be distinguished from all the others. For convenience they have been classified into metals and non-metals. We will now notice the

174. Distinguishing properties of the metals.

1. Metals have a peculiar brilliancy of surface, which is known as the **metallic lustre**.

2. As a rule, they are **good conductors of heat and electricity**. Metals are generally used for the plates of electric batteries, and for the conveyance of the current.

3. Metals have the **power of displacing the hydrogen of acids**, thus forming salts. (See Lesson XVII.)

4. Metals **cannot be dissolved without undergoing chemical change**. Thus, if zinc be dissolved in sulphuric acid, zinc sulphate is obtained on evaporating ; but, if sulphur be dissolved in carbon bi-sulphide, sulphur alone is left after evaporation.

5. All metals are fusible.

6. All are opaque.

7. All combine with oxygen to form oxides.

8. " " chlorine " chlorides.

9. " " sulphur " sulphides.

175. Classified list of the common elements.

Metals.		Non-metals.
Sodium	Tin	Hydrogen
Potassium	Manganese	Chlorine
Iron	Magnesium	Oxygen
Aluminium	Chromium	Nitrogen
Calcium	Silver	Carbon
Barium	Gold	Sulphur
Copper	Platinum	Phosphorus
Zinc	Mercury	Silicon
Lead		Iodine

All the metals above-mentioned are solids with one exception—mercury (quicksilver). Of the non-metals, the first four are gases, and the others are solids.

QUESTIONS ON LESSON XX.

1. What is an element? How many elementary substances are known?
2. What are the distinguishing characteristics of metals?
3. Write a list of about twenty-four elements, classifying them into metals and non-metals.

LESSON XXI.

IRON.

176. The physical properties of iron.

IRON is the most useful of all the metals. It is very tough. An iron wire only $\frac{1}{36}$ in. in diameter will support a weight of 80 lbs. When pure, it is granular or crystalline in structure, but, if rolled, becomes fibrous. Its melting point is very high; but it becomes soft at a much lower temperature; and, while in this condition, it may be welded, that is, pieces may be joined together by hammering. Iron may be drawn out to a very fine wire, and may be hammered or rolled out into a very thin sheet. The specific gravity of iron is 7.8.

177. The chief ores of iron.

Very little free iron is to be found in the earth's crust ; it is generally associated with other elements, as compound substances which are called the ores of the metal. The chief ores of iron are :—

(a) **Clay iron-stone.**—This is a carbonate of iron which is usually found associated with clay and lime, and is the most important source of iron in this country. The iron is separated

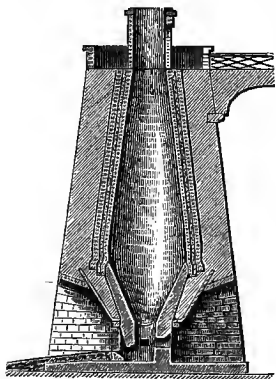


FIG. 57.—SECTION OF THE FURNACE USED FOR THE REDUCTION OF IRON ORE.

from this substance by heating it very strongly in a blast-furnace with coal and lime. The lime combines with the silica of the ore, forming a glassy substance which is called 'slag,' an oxide of iron is formed, and the coal unites with the oxygen, thus leaving the metal in a nearly pure state. As the iron is thus separated from the ore it runs to the bottom of the furnace in a molten condition, while the fused slag floats on its surface. When the process is complete, the iron is allowed to run out at openings in the bottom of the furnace into grooves made of sand,

where it solidifies. This iron is called in commerce 'pig-iron.' It is an impure cast iron, containing carbon, silicon, calcium, sulphur, and many other impurities.

(b) **Hematite.**—A native oxide of iron found very largely in this country. It is grey, black, brown, or red in colour, and often contains silica and alumina.

(c) **Iron pyrites.**—A sulphide of iron of a metallic brass-yellow or bronze-yellow colour. It is often found in cubical crystals, and sometimes presents a radiating structure. This ore is not worked for its iron, but large supplies of sulphur are obtained from it. It is used largely in the manufacture of sulphuric acid. The presence of sulphur in iron pyrites may be proved as follows :—

Experiment 116.—Powder a little of the ore, and heat it strongly in a hard glass tube held in a horizontal position so that one end may be kept cool. The vapour of sulphur passes off and condenses on the cool part of the tube. If the heat applied be sufficiently intense, a large proportion of the sulphur may thus be driven off, leaving a lower sulphide of iron which may be attracted by a magnet.

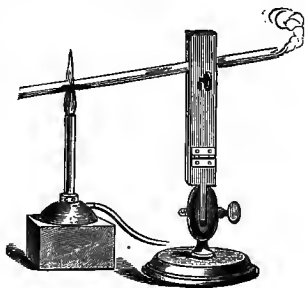


FIG. 58.—SHOWING THE METHOD OF SEPARATING SULPHUR FROM IRON PYRITES.

178. The varieties of iron.

Iron is never used for manufacturing purposes in a pure state. It is usually associated with carbon, silicon, and other impurities; and the properties of the iron may be greatly modified by the presence of these impurities, particularly the carbon. There are three varieties of iron—cast iron, wrought iron, and steel.

Cast iron is the most impure variety of iron. It contains from 5 to 6 per cent. of carbon, silicon, and other impurities. It is the first product of the iron ores in the blast-furnaces. Cast iron is granular or crystalline in structure, and is brittle. It cannot be welded or worked into shape by hammering. It is used largely for the manufacture of pipes and the heavier portions of machinery, and for various articles and utensils which are not likely to receive heavy blows.

Steel is the hardest kind of iron, and is consequently used for cutting instruments. It is purer than cast iron, containing only from 1 to 2 per cent. of carbon; but it is not so pure as wrought iron. Steel possesses great hardness and elasticity, but if heated strongly and allowed to cool slowly it loses both these properties, and becomes very much like wrought iron. It can again be hardened by raising it to a high temperature and then plunging into cold water or oil. This process is called **tempering**. By altering the rapidity of cooling every degree of hardness may be obtained. Steel may be rendered permanently magnetic. Wrought iron may be easily magnetised by

simply placing it near a magnet ; but it loses its magnetic properties as soon as the magnet is withdrawn. Steel is best magnetised by rubbing a magnet on its surface, and when treated in this manner it will retain its magnetic properties indefinitely.

Wrought iron, or **malleable iron**, is the purest and softest variety. It is very tough and fibrous, and is easily welded. Wrought iron is so called because it can be *wrought* into any required shape by hammering, cutting, and welding ; it is, therefore, largely used for making such articles as nails, horse-shoes, bolts, etc. Wrought iron is prepared from cast iron by the process called '*puddling*.'—The cast iron is melted in a furnace, and while the burning gases from coal are playing over its surface, it is kept in motion by means of the '*puddling-rod*.' The iron is thus deprived of the greater portion of its carbon.

179. Oxidation of iron.

When iron is exposed to moist air it becomes 'rusty' ; this is due to the formation of an oxide of iron, the oxygen being obtained from the air. When iron is strongly heated it oxidises more rapidly ; but in this case another oxide is produced in the form of black scales. This may be shown by heating a bar of iron to a white heat. After cooling, the black scales of the oxide may be removed from the bar by hammering. It will be remembered, too, (see Art. 64) that iron may be rapidly converted into a black oxide by burning in oxygen gas. Iron will not rust in perfectly dry air ; it will also retain its bright surface in pure water from which all air has been removed by boiling. The action of steam on iron was described in Art. 88.

180. The action of acids on iron.

Sulphuric, nitric, and hydrochloric acids all dissolve iron readily, forming iron sulphate, iron nitrate, and iron chloride respectively ; and, in each case, the salt may be obtained in a crystalline form by evaporation. It will be noticed, however, that when iron is thus dissolved a black deposit settles at the bottom. This is composed chiefly of the carbon which the iron contained, and would, consequently, be greatest when cast

iron is used. When iron is dissolved in sulphuric or hydrochloric acid, hydrogen gas is given off. The three salts of iron above-mentioned are all readily soluble in water.

181. Reduction of iron from the oxide.

Hydrogen and carbon are both valuable reducing agents; having a great affinity for oxygen gas, they are largely used for removing the oxygen from metallic oxides, thus reducing the metal to the pure state. On a large scale carbon (in the form of coal) is used for the reduction of iron; but the following simple experiment illustrates the reducing power of hydrogen.

Experiment 117.—Take a piece of hard ‘combustion tubing,’ about 9 inches long and $\frac{1}{2}$ inch or less in diameter, and place in the centre of it a little red oxide of iron. Now at each end fit a cork, perforated and provided with a glass tube passing through it. Attach one end to the delivery tube of the hydrogen apparatus (see Art. 82), so as to allow a current of *dry*¹ hydrogen to pass through the combustion tube. Apply heat to the iron oxide till red hot, and then allow the hydrogen to pass over. A change immediately commences —

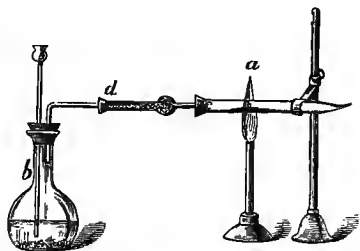


FIG. 59.—APPARATUS FOR THE REDUCTION OF IRON OXIDE BY MEANS OF HYDROGEN.

b, Flask in which hydrogen is being prepared; *d*, drying tube, containing small lumps of calcium chloride; *a*, combustion tube, containing the iron oxide.

the red oxide disappears, and black particles of iron remain in its place, of course with a loss of weight, for the oxygen of the iron oxide combines with the hydrogen passing over, and forms water-vapour, the condensed cloud of which may be seen issuing at the other end of the tube. When the action is over, and the tube cold, the iron particles may be shaken out. Being in a finely divided state, they will probably rapidly

¹ Hydrogen gas may be dried by passing it through a tube or bottle containing fragments of calcium chloride. Coal gas may be used instead of hydrogen.

oxidise as they pass through the air, becoming white hot through the rapid combustion.

182. Test for iron compounds.

Experiment 118.—Iron salts form a beautiful blue compound (Prussian blue) with ferrocyanide of potassium (yellow prussiate of potash). This substance is often used, therefore, as a test for the presence of iron in solution. To illustrate this, dissolve a little iron sulphate in water, render the solution slightly acid with a few drops of nitric acid, and then add a few drops of a solution of potassium ferrocyanide ; the result is a precipitate of Prussian blue.

183. Summary.

1. The chief ores of iron are :—

- (a) Clay iron stone, composed of iron carbonate and clay.
- (b) Hematite, composed of iron and oxygen.
- (c) Iron pyrites, composed of iron and sulphur.

2. All three kinds of iron (cast iron, wrought iron, and steel) contain carbon : cast iron contains most, and wrought iron least carbon.

3. Cast iron is crystalline and brittle, wrought iron is malleable and tough, steel is hard and brittle.

4. Iron oxidises slowly when exposed to moist air, and rapidly when heated in air or oxygen.

5. Iron may be obtained from its oxide by heating in hydrogen, or with carbon.

6. Iron compounds form Prussian blue with potassium ferrocyanide.

QUESTIONS ON LESSON XXI.

1. Describe the chief physical properties of iron.
2. What are the chief ores of iron ? Give the chemical composition of each.
3. Describe briefly the mode of preparing iron from clay iron-stone.
4. What is *slag* ? How is it formed ?
5. How would you prove that iron pyrites contains sulphur ?
6. Describe the chief differences between cast iron, wrought iron, and steel.

7. What is meant by the words 'tempering' and 'puddling'?
 8. What is iron rust? By what means could you prevent the rusting of iron?
 9. Describe the action of sulphuric, nitric, and hydrochloric acids on iron.
 10. How would you prepare crystals of iron sulphate?
 11. Describe a method of reducing iron from the oxide.
 12. How would you prove the presence of iron in a solution?
-

LESSON XXII.

LEAD.

184. The physical properties of lead.

LEAD is a bluish-white metal, and is so soft that it may be scratched with the finger nail. When pure it has that peculiar brilliancy of surface which is so characteristic of the metals, but when exposed to moist air it soon tarnishes by combining with oxygen. It does not oxidise in dry air; neither will it if placed in pure water which has been freed from air. Lead is malleable (*i.e.* may be extended by beating) and ductile (*i.e.* may be drawn out into wire). Being easily bent, it is largely used in making pipes for conveying gas and water. Lead easily melts (at 327° C.), and is largely used for casting bullets, etc. It is one of the heaviest metals, having a specific gravity of 11.3.

185. The ores of lead.

Pure lead is not found in nature; almost the whole of that used is obtained from a native sulphide of lead called **galena**. This ore of lead is found very largely in Cumberland, Cornwall, and other parts, associated generally with a small quantity of silver. The carbonate of lead (cerusite) is also found in nature. Black-lead, as has already been mentioned (Art. 90), is not a variety of lead, nor does it contain lead, but is one of the various forms of carbon.

186. The oxides of lead.

Experiment 119.—Melt some lead in an iron spoon or ladle, and remove the dross that floats on the surface by means of a piece

of iron plate. The metal now presents a bright silvery surface. Continue to apply heat to the metal, at the same time blowing a current of air over its surface with a pair of bellows. Some of the oxygen from the air will combine with the lead, and form a straw-coloured compound known as **litharge**. This oxide of lead, when strongly heated with silica, combines with that substance, forming a glass which is easily fusible, and therefore largely used in making the tubing so commonly employed in chemical experiments.

Experiment 120.—Place a little litharge in an iron spoon, and heat it to redness. Let it remain exposed in this heated condition for some time. It will gradually absorb oxygen from the air, and be changed into another oxide known as the **red oxide of lead, or red lead**. This oxide is also used in the manufacture of ‘soft’ glass.

187. To obtain lead from the oxide.

Experiment 121.—Mix a little lead oxide with powdered charcoal, and place the mixture in a hollow in a sound piece of charcoal. Now heat the mixture strongly by directing a blowpipe flame on it. After some time remove the remains of the mixture, and notice the little beads of metallic lead, which will produce black marks on paper when drawn over it. In this experiment the oxygen of the lead oxide combined with charcoal and passed off in the form of carbonic acid gas, leaving the lead in the metallic state.

188. Lead nitrate.

Experiment 122.—Pour some nitric acid into a large test-tube containing either some fragments of lead or a little lead oxide, and warm it over a spirit-lamp. The lead or oxide will be slowly dissolved by the acid, and, in the former case, red fumes will be given off as long as the action lasts. Allow the acid to dissolve as much as it is capable of doing, then set it aside to cool, and notice the deposit of crystals of lead nitrate. The crystallisation may be completed by the evaporation of the liquid. Dissolve some of this lead nitrate in water, and set aside the solution for future experiments.

189. Lead chloride.

Experiment 123.—This salt is not prepared by adding hydrochloric acid to lead or lead oxide, as the acid acts very slowly on these substances. It may be conveniently prepared by adding hydrochloric acid to a solution of lead nitrate; the lead chloride being thrown down as a white crystalline powder. Lead chloride is only slightly soluble in water, water dissolving only one-thirtieth its own weight of this substance.

190. Lead sulphate.

Sulphuric acid acts very slowly on lead, consequently the method of preparing lead sulphate is very similar to that of preparing the chloride.

Experiment 124.—Add sulphuric acid to a solution of lead nitrate, and lead sulphate is precipitated in the form of a white powder. Lead sulphate is insoluble in water.

191. Lead acetate, or sugar of lead.

This compound is a very useful salt of lead. It is soluble in water, and is prepared by dissolving lead in acetic acid. Lead acetate may be used instead of lead nitrate in the preparation of the sulphate and chloride.

192. Lead carbonate. White-lead.

Lead carbonate is insoluble in water, and is slowly formed when water containing carbonic acid gas comes in contact with lead.

White-lead is a combination of lead carbonate with lead oxide. It is largely used in making paints.

193. Summary.

1. Lead is a metallic element, soft, ductile, malleable, and easily fusible. S.G. = 11.3. Melting point = 327°C .
2. Lead oxidises in moist air, but not in dry air or in water freed from air.
3. The chief ores of lead are:—
 - (a) **Galena**—lead sulphide.
 - (b) **Cerussite**—lead carbonate.

4. Lead combines with oxygen, when heated, forming the yellow oxide (litharge) and red oxide (red lead).

5. Lead may be obtained from lead oxide by strongly heating with charcoal.

6. The following table gives a summary of the chief facts concerning the more important compounds of lead :—

Chemical name	Common name	Composition	In water	Preparation	All poisonous
Lead oxide	Litharge	Lead and oxygen	Insoluble	Heat lead in air	
Lead oxide	Red lead	Lead and oxygen	Insoluble	Heat litharge in air	
Lead nitrate	—	Lead, nitrogen, and oxygen	Soluble	Dissolve lead in nitric acid	
Lead chloride	—	Lead and chlorine	Slightly soluble	Add hydrochloric acid to solution of nitrate	
Lead sulphate	—	Lead, sulphur, and oxygen	Insoluble	Add sulphuric acid to solution of nitrate	
Lead acetate	Sugar of lead	Lead, carbon, and oxygen	Soluble	Dissolve lead in acetic acid	
Lead carbonate	White lead	Lead, carbon, oxygen	Insoluble		

QUESTIONS ON LESSON XXII.

- Describe the physical properties of lead.
- What are the chief ores of lead? Describe the composition of each.
- What are the chief compounds of lead and oxygen? Describe their preparation.
- What is red lead? How is it prepared?
- How would you obtain lead from lead oxide?
- Describe the preparation and properties of lead nitrate.
- How would you prepare lead chloride?
- How would you prepare lead sulphate?
- What is sugar of lead? How is it prepared? What are its properties?
- What is white-lead?
- Describe two methods of preparing lead sulphate and lead chloride.
- Give a list of the chief compounds of lead, and describe the composition of each.

LESSON XXIII.

COPPER.

194. Physical properties of copper.

COPPER may be distinguished from all other metals by its red colour. It is a little heavier than iron, its specific gravity being 8.8. It is very ductile and very malleable. It is very tenacious; a wire of only one-fifteenth of an inch in diameter will support a weight of $2\frac{1}{2}$ cwts. It is one of the best conductors of heat and electricity. If heated to a bright red heat it melts, and, at a white heat, it gives off a vapour which burns with a green flame. A fine copper wire may be fused in the flame of a Bunsen burner; but an iron wire, in the same flame, would simply be softened.

195. Occurrence in nature.

Native copper is found largely in Cornwall and Siberia, but our supplies of this metal are obtained chiefly from an ore containing copper, iron, and sulphur, called **copper pyrites**. This mineral also is very abundant in Cornwall.

196. Oxidation of copper.

Copper undergoes no change in dry air, nor does it oxidise under any conditions at ordinary temperatures. When copper is exposed to moist air it becomes covered with a green crust.

Experiment 125.—When heated strongly in air, copper combines with oxygen, forming a black oxide. This may be illustrated by heating copper turnings in a Bunsen burner.

197. Reduction of copper oxide by hydrogen.

Experiment 126.—Copper may be obtained from its oxide by an experiment similar to that described in Art. 181. Pass a current of dry hydrogen or of coal gas through a tube containing copper oxide heated to a red heat. The hydrogen combines

with the oxygen, forming water vapour, and, after the action is complete, the tube will contain pure red copper in the place of the black oxide.

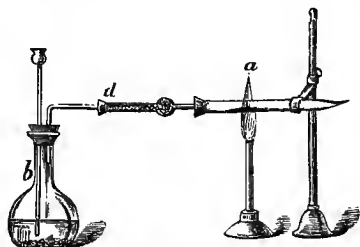


FIG. 60.—APPARATUS ILLUSTRATING EXPERIMENT 126.

a, tube containing copper oxide; *b*, flask in which hydrogen gas is being prepared; *d*, drying tube, containing small lumps of calcium chloride.

198. Action of acids on copper.

Sulphuric and hydrochloric acids do not dissolve copper at ordinary temperatures. Heated sulphuric acid dissolves it rapidly, converting it into **copper sulphate** or blue vitriol, and giving off sulphur dioxide—the same gas as that produced by the

combustion of sulphur in air or oxygen. (See Art. 62.)

Experiment 127.—Add *cold* nitric acid to copper. The metal is attacked with great energy. A colourless gas (nitric oxide)

is given off, which immediately on coming in contact with the air combines with oxygen and forms a reddish-brown gas. **Copper nitrate** is also formed, and may be obtained from the solution (by slow evaporation) in the form of dark greenish-blue crystals. Nitric oxide gas may be collected in a jar standing over water

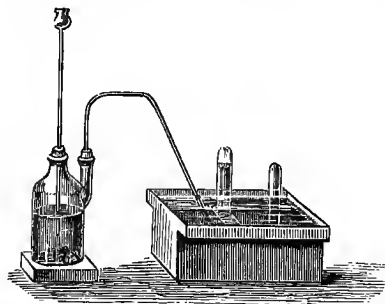


FIG. 61.—APPARATUS FOR PREPARING AND COLLECTING NITRIC OXIDE BY THE ACTION OF NITRIC ACID ON COPPER.

by using an apparatus similar to that shown in the accompanying sketch.

199. Verdigris.

When copper is allowed to remain for some time in contact with organic matter, it frequently becomes covered with a light

green substance called verdigris. This is **copper acetate**, and is formed by the action of the acetic acid in the organic substance.

200. The alloys of copper.

The properties of metals are often modified to a considerable extent by mixing with other metals, and the mixtures thus produced are frequently especially adapted for certain purposes for which the single metals would be inadequate. Such mixtures of metals are called **alloys**, and are prepared by thoroughly mixing the constituents while in a molten state. The alloys of copper are very useful, and extensively used. The most important are—

Brass.—Composed of copper and zinc in the proportion of two-thirds or three-fourths of the former, and one-third or one-fourth of the latter, with frequently a little lead. This alloy is of a fine yellow colour, resembling that of gold, but a little paler. Brass is much harder than copper, and works much more easily with the lathe ; it is also very ductile and malleable.

Tinsel consists of the same metals as brass in different proportions. It may be hammered into extremely thin plates.

Bronze.—There are many varieties of bronze used for different purposes, all composed of copper and tin in various proportions.

Bell-metal.—Composed of from three to six parts of copper and one of tin.

German silver.—An alloy of zinc, nickel, and copper.

Gun-metal.—Composed of nine parts of copper and one of tin.

The **Bronze used for English coinage** consists of 95 parts of copper, 4 parts of tin, and 1 of zinc.

The silver used in coining money is always alloyed with a small quantity of copper, which increases its hardness and resistance to wear. Gold, which is softer than silver, is also hardened by the admixture of a little copper or silver, or sometimes both.

201. The displacement of copper from its solution.

Experiment 128.—Prepare a solution of copper sulphate, and add a drop of sulphuric acid. Introduce into this solution a piece of bright and clean steel—a knife-blade or a key. In a very short time the steel will be covered with a deposit of the red metallic copper which has been displaced from its combination with sulphur and oxygen.

202. Summary.

1. Copper is a metallic elementary substance, red, malleable, ductile, tenacious, and a good conductor of heat and electricity.
2. Copper is found native. Its chief ore is copper pyrites.
3. It does not oxidise unless strongly heated.
4. Copper oxide may be reduced by heating in hydrogen.
5. Hydrochloric acid has no action on copper.
6. Heated sulphuric acid dissolves copper, forming copper sulphate and sulphur dioxide.
7. Nitric acid dissolves copper, forming copper nitrate.
8. Verdigris is copper acetate.
9. Copper is hardened by the admixture of zinc and tin.

QUESTIONS ON LESSON XXIII.

1. Describe the chief physical properties of copper.
2. Compare the properties of iron, lead, and copper.
3. What is meant by *native* copper?
4. What is the most abundant ore of copper? Describe its chemical composition.
5. Under what conditions will copper combine with oxygen?
6. What chemical action takes place when copper is exposed to the atmosphere?
7. How would you obtain pure copper from copper oxide?
8. What is the action of hydrochloric acid on copper?
9. Describe fully the action of sulphuric acid on copper. Give the composition and properties of the products of this action.
10. How would you prepare crystals of copper sulphate?
11. What is blue vitriol?
12. Describe fully the action of nitric acid on copper.
13. How would you prepare crystals of copper nitrate?
14. What is verdigris? How is it formed?
15. What is an alloy? Mention the chief alloys of copper.
16. How would you prove that the blue salt called blue vitriol contains the red metal, copper?

LESSON XXIV

MERCURY.

203. The physical properties of mercury

MERCURY is the only metal liquid at ordinary temperatures, and is the heaviest liquid known. It will balance, on a pair of scales, $13\frac{1}{2}$ times its own volume of water—its specific gravity is 13.6. If reduced to a temperature of 40° C. below zero (-40° C.), mercury freezes, and is then a malleable metal with a crystalline structure. Mercury boils at 350° C., giving off a colourless vapour which condenses readily on cooling.

Experiment 129.—This may be illustrated by heating about an ounce of mercury in a test-tube over the flame of a spirit-lamp or gas-burner, holding the tube in an oblique position. The vapour given off condenses on the cool part of the tube in the form of little globules.

Mercury is often purified by the process called distillation ; for, like water, when converted into vapour, it leaves behind most of its impurities. It is used largely in the manufacture of thermometers, as it expands and contracts regularly, and does not freeze except in countries where the cold is intense. Being the heaviest liquid known, it is used in making barometers. The atmosphere will support a vertical column of mercury about 30 inches high, and the varying pressure of the air is measured by the height of the column of mercury supported. A barometer containing water instead of mercury would be more than 30 feet in height, since 30 feet of water presses no more than 30 inches of mercury.

204. Occurrence in nature.

Mercury is found *native* in Spain, California, and other countries, but our chief supplies are obtained from an ore called **cinnabar**. This mineral is a sulphide of mercury, having the same composition as the well-known pigment vermilion.

205. Oxidation of mercury.

Mercury does not tarnish in either dry or moist air, but always retains its bright metallic surface. On this account it is classed among the *noble* metals—the metals that do not tarnish

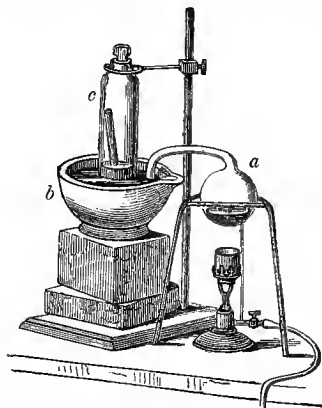


FIG. 62.—EXPERIMENT FOR ILLUSTRATING THE CHEMICAL COMPOSITION OF AIR BY THE OXIDATION OF MERCURY.

a, vessel containing mercury to be oxidised, provided with a bent neck which communicates with a closed vessel (*c*) of air; *b*, vessel containing mercury, which closes the mouth of the vessel *c*. The mercury slowly rises in *c* as the air is deprived of its oxygen.

by exposure to air. However, if mercury be exposed to the air for some days, and kept at a temperature of about 300°C ., it slowly combines with oxygen, and forms a red oxide, commonly known as the **red oxide of mercury**, or **red precipitate**. If this experiment be conducted in a closed vessel, with a limited supply of air, it will be noticed that the air slowly decreases in volume, owing to the removal of its oxygen—nitrogen only remaining. It was in this manner that Lavoisier first discovered the chemical composition of air. The oxide of mercury thus formed, if more strongly heated, is again decomposed

into its elements mercury and oxygen, as described in Exp. 29.

206. Action of acids on mercury.

Hydrochloric acid does not act on mercury. Sulphuric acid, when heated, dissolves the metal, forming mercury sulphate, and liberating sulphur dioxide. Nitric acid dissolves mercury rapidly, forming mercury nitrate, and setting free the colourless gas which gives rise to reddish-brown fumes in contact with air. The student should compare the action of acids on mercury, with their action on copper as described in Art. 198.

207. Amalgams.

Mercury has the power of dissolving some metals, just as water dissolves sugar, thus forming alloys which may be liquids, pastes, or solids, according to the proportion of mercury present in the mixture. The alloys of mercury are termed *amalgams*, and metals which have been wholly or partially dissolved by means of mercury are said to have been *amalgamated*.

Experiment 130.—Stretch a piece of tinfoil over the top of a vessel, and make its upper surface slightly concave. Pour on this tinfoil a small quantity of mercury. The mercury will soon dissolve the tin beneath it and fall through with the liquefied tin to the bottom of the vessel.

Experiment 131.—Place some tinfoil and mercury in a vessel together. After a time all the tin will have been dissolved, and a *tin amalgam* will have been formed.

Gold is separated from the quartz with which it is found associated in nature, by dissolving it out with mercury, and afterwards heating the amalgam to drive off the mercury as vapour. Silver, zinc, copper, and many other metals may also be amalgamated with mercury.

208. The displacement of mercury from its solutions.

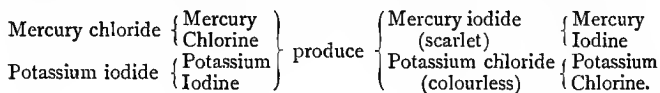
Experiment 132.—Dip a piece of bright iron or copper into a solution of some soluble salt of mercury (mercury nitrate, for example). The metal will soon be covered with a deposit of mercury in the form of a grey powder, which becomes bright when rubbed.

209. Separation of mercury from its compounds.

Experiment 133.—Mix a little powdered mercury chloride with dry sodium carbonate, introduce the mixture into a hard glass tube and apply heat. The mercury will soon be liberated from the chlorine, pass off as a vapour, and condense in the form of little globules on the cool part of the tube. In the same manner mercury may be separated from any of its compounds.

210. An example of the mutual exchange of elements.

Experiment 134.—The following interesting experiment illustrates a mutual exchange of elements between two compounds:—Make weak solutions of mercury chloride (corrosive sublimate) and potassium iodide. Add a little of the latter solution to the former, and a precipitate will be formed which is at first a pale yellow, but speedily changes to a salmon colour and then to scarlet. The mercury chloride contains the elements mercury and chlorine; the potassium iodide contains the elements potassium and iodine. The former compound gives up its chlorine for iodine, while the latter gives up its iodine for the chlorine. The coloured compound formed is the mercury iodide produced by this mutual interchange.

**211. Summary.**

1. Mercury is the only metal liquid at ordinary temperatures, is the heaviest liquid known, and has a specific gravity of 13.6.

2. It boils at 350° C., and freezes at -40° C. When it boils it gives off a colourless vapour, and when solid it is crystalline and malleable.

3. Mercury is found native. The chief ore of mercury is cinnabar, composed of mercury and sulphur.

4. Mercury does not tarnish when exposed to air, but oxidises slowly when heated to near its boiling point.

5. Hydrochloric acid does not attack mercury. Hot sulphuric acid dissolves it, forming the sulphate. Nitric acid dissolves mercury, forming mercury nitrate.

6. Mercury dissolves many metals, forming amalgams.

7. Mercury is displaced from its solutions by bright iron or copper.

8. It may be separated from any of its compounds by heating with sodium carbonate.

QUESTIONS ON LESSON XXIV.

1. Describe the physical properties of mercury.
 2. Why is mercury used for thermometers and barometers?
 3. What changes take place when mercury is heated gradually to a high temperature?
 4. What are the boiling and freezing points of mercury?
 5. In what forms does mercury exist in nature?
 6. What is red precipitate? How may it be formed?
 7. Describe the action of heat on red oxide of mercury.
 8. What is an amalgam? Give the names of some metals that will amalgamate with mercury.
 9. How could you prove the presence of mercury in a solution?
 10. What is cinnabar? How would you prove that it contains mercury?
 11. How would you obtain mercury from mercury chloride?
-

LESSON XXV.

SODIUM.

212. The physical properties of sodium.

SODIUM is a very soft metal—so soft that it can be worked into any shape by the pressure of the fingers. If cut or broken, the parts may be easily welded together again.

Experiment 135.—Throw a piece of sodium on water. The metal swims about on the surface, decomposing the water, and gradually disappearing. Some of the hydrogen of the water is set free, and the sodium is converted into caustic soda. It will be noticed, too, that, while the sodium is swimming on the water, it is always globular in form. This is due to the fact that it has been melted by the heat consequent on the chemical action. It will thus be seen that the fusion point of the metal is very low.

213. Occurrence in nature.

Sodium is never found native, its affinity for oxygen and other elements being so great. Common salt, which is so largely distributed in the sea, as well as in the rocks as **rock salt**, contains more than one-third of its weight of sodium. Most natural waters contain sodium in combination with other

elements, chiefly in the form of the chloride and sulphate. The metal sodium is prepared chiefly from the carbonate, by heating it to whiteness in an iron retort with powdered charcoal.

214. Oxidation of sodium.

Experiment 136.—Cut a piece of sodium with a knife; it presents a bright silver-white surface. But this brilliancy lasts only for a very short time. The metal is almost immediately covered with a white film. If the air is perfectly dry this film consists of sodium oxide; but if any moisture is present in the air, the metal combines with the hydrogen of the water vapour as well as with oxygen, thus producing sodium hydrate or caustic soda. The affinity of sodium for the elements of water is so great that it is difficult to preserve the metal from oxidation, even in a stoppered bottle, unless some means be taken to exclude the air effectually. The plan usually adopted is that of covering the sodium with a liquid such as naphtha, which contains no oxygen, and which has a lower specific gravity than the sodium, thus allowing the metal to sink.

215. The action of sodium on water.

This action has already been briefly noticed; but the following experiment is a further illustration of it.

Experiment 137.—Invert a test-tube full of water in a dish of water. Now take a piece of sodium, about the size of a large pea, and wrap it up closely in a piece of blotting-paper or fine gauze. Pass the sodium quickly into the inverted tube; hydrogen will be liberated, and will drive out the water from the tube. Remove the tube and test the gas by observing its properties. Next add a little reddened litmus to the water in the dish. The litmus will turn blue, showing the presence of an alkali. This alkali is caustic soda. (See Arts. 151 and 152.)

216. The chief compounds of sodium.

The principal compounds of sodium are—

(1) **Common salt, or sodium chloride.** This is a binary compound, containing sodium and chlorine—39·3 per cent. of the former to 60·7 per cent. of the latter. In Art. 135 the

method of obtaining the chlorine from common salt is described. **Rock salt**, found so largely in the earth's crust, is sodium chloride. **Sea salt** is composed chiefly of this same compound. (See Art. 28.)

(2) **Common washing soda** is a **sodium carbonate** ; that is, a combination of sodium with carbon and oxygen. Carbonic acid gas may, therefore, be driven from it by the addition of an acid.

(3) **Bicarbonate of soda** is very similar in composition to the last-named compound—it is a sodium carbonate containing hydrogen.

217. Summary.

1. Sodium is very soft, easily fusible, and lighter than water.

2. It has a silver-white surface when freshly cut, but soon oxidises.

3. It occurs in nature in rock salt and sea salt. Sodium sulphate is also found in natural waters.

4. When sodium is exposed to dry air it forms the oxide ; if the air contains water vapour, the hydrate (caustic soda) is produced.

5. Sodium decomposes water, setting free hydrogen, and forming caustic soda.

6. The chief compounds of sodium are common salt, washing soda, and bicarbonate of soda.

QUESTIONS ON LESSON XXV.

1. What are the chief physical properties of the metal sodium ?
 2. In what forms is sodium found in nature ?
 3. Name some important compounds of sodium, and give the composition of each.
 4. What chemical change takes place when sodium is exposed to (a) dry air, (b) moist air ?
 5. Describe fully the action of sodium on water.
 6. How would you prepare hydrogen from water by means of sodium ?
 7. What is caustic soda ? How may it be produced from sodium ?
 8. Describe the composition of 'washing soda.' How would you prove that it is a carbonate ?
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LESSON XXVI.

OTHER USEFUL METALS.

IN the present lesson we shall learn the properties and uses of other important metals not previously described.

218. Zinc.

Zinc is a bluish-white metal, having a specific gravity of 6.9. It is somewhat brittle at ordinary temperatures, but when heated slightly it may be hammered or rolled out into a sheet. The melting temperature of zinc is 423° C. (not very much higher than that of lead), so that it may be melted in an ordinary fire. If heated strongly in the presence of air, zinc burns with a bright greenish-white flame, forming zinc oxide. This may be illustrated by holding a strip of thin zinc foil in the flame of a Bunsen burner. As zinc is not acted on by air or water, it is used largely for covering the roofs of houses, and for lining water cisterns. It is also used to form a protecting covering for iron goods—the so-called **galvanised iron** being sheet iron covered with a thin coating of zinc. Dilute acids dissolve zinc readily, giving off hydrogen gas.

219. Tin.

This is a silver-white metal, and is soft, malleable, and ductile. Its specific gravity is 7.3. Its melting temperature (235° C.) is considerably lower than that of lead, and a piece of the metal may be easily melted in the flame of a candle. The '**tin-plate**' which is used so largely for making kettles, cans, and other utensils, is really sheet iron, covered with a thin coating of tin; the latter metal, not being oxidised by air or water, serves to protect the iron. Hydrochloric acid dissolves tin, forming a chloride of the metal, and liberating hydrogen. The substance called **putty powder** is an oxide of tin, which may be prepared by burning the metal.

220. Magnesium.

This metal closely resembles zinc both in appearance and chemical properties. It melts at a dull red heat, and boils when heated to bright redness. Magnesium does not oxidise in dry air at ordinary temperatures ; but, when heated strongly in air, it burns with a dazzling white light, forming an oxide which is called *magnesia*. Magnesium rapidly dissolves in hydrochloric and sulphuric acids, and in each case hydrogen gas is evolved. The substance known as Epsom salts is magnesium sulphate. Magnesium is one of the lightest of metals, its specific gravity being only 1.75.

221. Silver.

Silver is a very ductile metal, and is the best conductor of heat and electricity known. It is readily dissolved by nitric acid, silver nitrate being formed, and nitric oxide evolved. Silver is not oxidised in air, but readily combines with sulphur, forming a black sulphide. Coal gas contains small quantities of sulphuretted hydrogen, and, when this gas comes in contact with silver, it gives up its sulphur, which combines with the metal. It is for this reason that silversmiths burn their gas chiefly outside their shops, and also protect their silver goods by enclosing them in air-tight cases. If a silver spoon be used in eating an egg it is blackened by the combination of the metal with sulphur from the egg. Silver is sometimes used in a pure state, but for coinage, and for the manufacture of silver plate, it is generally mixed with a small amount of copper. The specific gravity of silver is 10.5.

222. Gold.

Gold is always found in the free or uncombined state. It is one of the heaviest of metals, having a specific gravity of 19.34. It is extremely malleable and ductile, and in its pure state is nearly as soft as lead. Gold does not oxidise at any temperature, either in dry or moist air, hence it is used largely in the manufacture of ornaments. Single acids do not act on gold, but the metal may be dissolved in *aqua regia* (a mixture of nitric and hydrochloric acids) when gold chloride will be

formed. Gold, being too soft to be used alone, is generally alloyed with copper. **Pure gold** is represented by the number 24, and is described as 24-carat gold. The **guinea gold**, or **standard gold** of our country, is 22-carat gold, that is, an alloy of 22 parts gold and 2 parts copper: this is the quality used for our coinage; but 18-carat gold is the best for manufacturing purposes.

223. Platinum.

Platinum is a silver-white metal which is heavier than gold, its specific gravity being 21.5. It is also much harder than gold, and is quite infusible by any ordinary furnace heat. In its chemical properties platinum closely resembles gold; it will not oxidise in air at any temperature; is not soluble in single acids; and dissolves in *aqua regia*, forming platinum chloride.

224. Potassium—Calcium—Barium.

These metals are very abundant in the earth's crust, but are never found in a free state; in fact, they have such a powerful affinity for oxygen and other elements that it is difficult to separate them from their compounds. They all resemble sodium in chemical properties.

QUESTIONS ON LESSON XXVI.

1. Describe the chief properties of the metal zinc.
 2. What is *galvanised iron*? Why is iron 'galvanised'?
 3. What are the chief properties of tin? What is '*tin-plate*'? What are its advantages over iron plate?
 4. Describe the metal magnesium.
 5. What is magnesia? How may it be prepared?
 6. What are the chief properties of silver? What is the cause of the blackening of this metal when exposed to air?
 7. What is the action of nitric acid on silver?
 8. Why is gold so largely used in the manufacture of ornaments?
 9. What liquid would you use to dissolve gold? Describe the action of this liquid on the metal.
 10. What is '*guinea gold*'? Why is gold generally alloyed with copper?
 11. Describe the chief properties of platinum.
 12. What do you know about potassium, calcium, and barium?
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PART II.—ORGANIC CHEMISTRY.

LESSON XXVII.

ORGANIC COMPOUNDS.

225. The composition of plants and animals.

OF the many elements contained in the earth's crust, only a few exist to any extent in animal and vegetable substances. It may be said that organic substances are composed almost entirely of the four elements carbon, hydrogen, oxygen, and nitrogen. Nitrogen is often absent, and sometimes also oxygen ; but carbon is such an essential ingredient of animal and vegetable compounds, that organic compounds are often spoken of as the **carbon compounds** ; and **organic chemistry** is often said to be the study of the carbon compounds. But it must be remembered that many inorganic or mineral substances contain the element carbon ; e.g. all the carbonates. Also, there are some compounds that are to be found both in the mineral kingdom and in animals and plants, and therefore belong both to inorganic and organic chemistry. Organic chemistry may, however, be defined as the study of the carbon compounds derived from animal and vegetable substances. The carbon compounds are extremely numerous, and of varied character. Some are acids, some are salts, and others are neutral bodies. The chief of them will be described in the following lessons.

226. Test for organic compounds.

When an organic substance is heated in a glass tube, it is decomposed into compounds of a more simple nature. These compounds are chiefly gases, similar to those gases which are evolved during the ordinary decay of organic matter. In

addition to these are vapours which condense on the cool part of the tube. This destruction of an organic substance by heat is generally accompanied by an odour resembling that of coal tar, and a residue of carbon is generally left behind. Many organic compounds are very similar in appearance to certain mineral substances, but we may distinguish between them by using the above test.

227. Organic substances generally contain a certain amount of inorganic matter.

Experiment 138.—Take a small piece of an animal or vegetable substance, such as bread, meat, wood, or cotton, and heat it in a hard glass tube. It is decomposed, as described above, leaving a black residue consisting chiefly of carbon. Now heat this residue very strongly on a piece of platinum foil until all the carbon has been converted into carbonic acid gas. A *white* or *whitish* substance remains, called the **ash**. This ash is the inorganic or mineral matter which the organic substance contained. A similar change may be noticed in an ordinary fire—after all the carbon has been burnt up, a whitish ash remains. Thus we have a simple means of calculating the amount of mineral matter derived from the soil by plants; also the amount of inorganic matter present in the bodies or parts of animals.

228. The chief differences between organic and inorganic substances.

There are some interesting points of difference between organic and mineral compounds. In the first place, we find that all the known elements enter into the composition of mineral compounds, while only about twenty are to be found in organic substances; and of these twenty only four are found in very large quantities. Again, organic substances may generally be distinguished from minerals by their natural tendency to decay; while nearly all the mineral substances may be exposed for any length of time without undergoing decomposition.

229. Summary.

1. Organic chemistry is the study of the carbon compounds found in animals and vegetables.
2. These organic compounds consist chiefly of the elements carbon, hydrogen, oxygen, and nitrogen.
3. Organic compounds generally leave a residue of carbon when heated.
4. When burnt, most of them leave behind a whitish ash, composed of mineral matter.
5. They generally undergo decomposition when exposed to air.

QUESTIONS ON LESSON XXVII.

1. What is meant by the terms *organic* and *inorganic*?
 2. What do you know of the composition of organic substances?
 3. By what means would you distinguish between organic and inorganic compounds?
 4. What is ash? How would you determine the amount of vegetable substance and the amount of mineral matter in a piece of paper?
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LESSON XXVIII.**ACETIC ACID.****230. General properties of acetic acid.**

ACETIC acid is one of the most important organic acids. It is a colourless liquid, solidifying when pure at about $15^{\circ}\text{C}.$, forming large transparent crystals, and boiling at $120^{\circ}\text{C}.$ It may be easily distinguished by its very powerful, pungent odour. **Vinegar** is an impure variety of this acid; and the ordinary colourless acetic acid of the shops, which is sometimes called **white vinegar**, is the pure acetic acid diluted with water, the acid being readily soluble in water. Acetic acid possesses all the general properties of acids as mentioned in Art. 148. It turns blue litmus red, has a sharp, sour taste, neutralises alkalies, and decomposes carbonates.

231. Action of acetic acid on carbonates. Acetates.

When added to a carbonate, acetic acid sets free the carbonic acid, and combines with the metal, forming the acetate of the metal. The following experiment illustrates this action.

Experiment 139.—Put a little sodium carbonate into a test-glass, and add acetic acid very slowly and gradually till the gas ceases to be evolved. A lighted taper plunged into this gas is instantly extinguished ; and it may be proved to be carbonic acid gas by passing a little of it into lime-water, when a precipitate of chalk will be formed. Now evaporate the liquid to dryness, and a white crystalline substance will remain. This is the salt known as sodium acetate. Other acetates may be formed in a similar manner by using the carbonates containing other metals.

232. Action of acetic acid on alkalies.

Experiment 140.—All alkalies may be neutralised by acetic acid, and in each case an acetate will be formed. Prepare a small quantity of a strong solution of potash, and add to it sufficient solution of litmus to give it a distinctly blue colour. (It will be remembered that the *alkali* has no action on *blue* litmus.) Then add acetic acid, drop by drop, stirring all the time, till the liquid *just begins to turn red*, and remains permanently of a purplish colour. The solution is now *neutral*, the acid and the alkali have *exactly* neutralised each other, and formed a salt. Evaporate the liquid to obtain the salt in a crystalline form. This salt is **potassium acetate**. Other acetates may be similarly produced by neutralising the acid with the alkalies of other metals.

233. Composition of acetic acid.

Acetic acid contains the three elements carbon, hydrogen, and oxygen. Its percentage composition, by weight, is carbon 40 parts, hydrogen 6·7 parts, and oxygen 53·3 parts.

234. The formation of acetic acid.

The impure variety of acetic acid known as vinegar is largely prepared by the exposure of the poorer wines and beers to the air. These liquors contain alcohol, and the alcohol combines with oxygen of the air, thus changing into acetic acid. It has been found that this change cannot take place without the presence of a minute living organism, which, during its rapid growth and multiplication, acts as a medium by which the oxygen is transferred from the air to the alcohol. This organism is really a minute vegetable—a kind of *fungus*—which floats on the liquid, and has received the name *mycoderma aceti*.

Acetic acid is also prepared by the **destructive distillation of wood**. Wood is heated in a large retort, and the various vapours which distil over are condensed in a cool receiver. Among these vapours is that of acetic acid, and consequently the liquefied acid is found mixed with the other products of distillation in the receiver, and is afterwards separated from them.

235. Action of acetic acid on iron.

Experiment 141.—Pour some strong acetic acid on iron borings, or scraps of thin iron wire. The iron is dissolved, and hydrogen gas is evolved. Filter the liquid when all action has ceased, and the substance thus obtained is a solution of **iron acetate**.

236. Lead acetate.

Experiment 142.—Dissolve a little lead oxide in acetic acid, and evaporate the liquid remaining to obtain the white crystals of lead acetate. This salt is soluble in water, and has a sweetish taste. It is commonly known as **sugar of lead**.

237. Sodium acetate.

This salt is also soluble in water. Like the other acetates, it does not possess the odour of acetic acid. But if warmed in a test-tube with a little dilute sulphuric acid, the vapour of acetic acid is given off, and may be recognised by its odour. It may also be collected in a cooled receiver.

238. Summary.

1. Acetic acid is a vegetable acid containing the elements carbon, hydrogen, and oxygen.
2. Vinegar is a variety of impure acetic acid. White vinegar is diluted pure acetic acid.
3. Acetic acid decomposes carbonates, setting free carbonic acid gas, and forming acetates.
4. Acetic acid neutralises alkalis, forming the acetates of the metals present.
5. The alcohol in wines and beers is converted into acetic acid by the absorption of oxygen.
6. Acetic acid is one of the products of the destructive distillation of wood.
7. Acetic acid dissolves iron, forming iron acetate, and liberating hydrogen gas.
8. Sugar of lead may be produced by dissolving lead oxide in acetic acid.
9. Acetates do not possess the odour of acetic acid, but the acid is given off as a vapour when acetates are heated with a stronger acid.

QUESTIONS ON LESSON XXVIII.

1. Describe the chief properties of acetic acid.
 2. What is vinegar? How is it made?
 3. Describe three experiments by which you would prove that vinegar is an acid substance.
 4. Describe fully the changes produced when acetic acid is added to potassium carbonate.
 5. How would you prepare sodium acetate by means of caustic soda and acetic acid?
 6. Give the percentage composition of acetic acid.
 7. Describe the changes which take place when wine or beer is exposed to air.
 8. How may acetic acid be prepared from wood?
 9. Describe the action of acetic acid on iron.
 10. What is *sugar of lead*? How may it be prepared?
 11. How would you obtain acetic acid from an acetate?
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LESSON XXIX

TARTARIC ACID.

239. The properties of tartaric acid.

THIS is another vegetable acid. It occurs in the juices of many fruits, especially grapes. Tartaric acid is a crystalline solid, possessing no odour, and is readily soluble in water. It has all the properties common to acids. Its solution reddens litmus, and decomposes carbonates.

240. Composition.

Tartaric acid is composed of the same elements as acetic acid, but in different proportions, viz. carbon 32 parts, hydrogen 4 parts, and oxygen 64 parts.

241. Preparation of tartaric acid.

This acid is obtained chiefly from the juice of the grape. The juice, after being pressed from the fruit, is allowed to ferment. This fermentation is a chemical action, giving rise to an impure potassium tartrate, named in commerce 'argol,' from which the acid is prepared by the separation of the metal potassium.

242. Action of tartaric acid on carbonates. Tartrates.

Experiment 143.—Dissolve a few crystals of tartaric acid in water, and add, a little at a time, a solution of washing soda. (It will be remembered that common washing soda is sodium carbonate.) Carbonic acid gas is rapidly given off, and the action continues till the whole of the acid has been neutralised. When this is the case, evaporate the liquid, and a crystalline mass of **sodium tartrate** is produced. Other tartrates may be similarly prepared by decomposing various carbonates.

243. Effervescing draughts.

Many effervescing drinks are made by dissolving tartaric acid with sodium bicarbonate, with the addition sometimes of sugar or other substances to render the whole palatable.

Experiment 144.—Mix about a dram of sodium bicarbonate with three-quarters of a dram of powdered tartaric acid. The two compounds will not combine while in the solid state. Now add a few ounces of water, and effervescence is at once set up as the substances dissolve. The effervescence is due to the rapid production of carbonic acid gas. The above mixture, if rendered pleasant to the taste by the addition of ground loaf-sugar and essence of lemon, corresponds with the beverage commonly known as sherbet or lemon kali. The cooling effect of such drinks is due to the rapid liberation of the gas.

Seidlitz powders are contained in two separate papers. One contains 40 grains of sodium bicarbonate and 120 grains of Rochelle salt; the other 37 grains of tartaric acid. The contents of these papers are dissolved in separate glasses, and, on mixing the solutions, the acid rapidly sets free carbonic acid gas from the carbonate.

244. Cream of tartar.

This well-known salt is a potassium tartrate. It is prepared from 'argol,' which is dissolved in water, and filtered through crushed charcoal and pipe-clay to remove all colouring matter. The liquid is then evaporated till the colourless crystals of cream of tartar are deposited. Cream of tartar has a sour taste, reddens litmus, and is only slightly soluble in cold water. It is not an acid, but one of those salts in which the acid is only partially neutralised by the base.

245. Summary.

1. Tartaric acid occurs in fruits. It contains carbon, hydrogen, and oxygen.
2. It is prepared chiefly from argol—the impure potassium tartrate obtained from the juice of the grape.

3. Tartaric acid decomposes carbonates, setting free carbonic acid gas, and forming tartrates.

4. Many effervescing drinks contain tartaric acid and sodium bicarbonate.

5. Cream of tartar is a potassium tartrate, possessing acid properties.

QUESTIONS ON LESSON XXIX.

1. What are the chief properties of tartaric acid ?
2. Describe the chemical composition of tartaric acid.
3. Whence do we obtain our chief supplies of this acid ?
4. Describe fully the action of tartaric acid on sodium carbonate.
5. What is cream of tartar ? Whence is it obtained ?
6. How would you prove that tartaric acid contains carbon ?
7. What is the cause of the effervescence produced by dissolving sherbet or seidlitz powders in water ?

LESSON XXX.

FATS AND OILS.

246. The general properties of fats and oils.

FATS and oils are all neutral bodies, being composed of an acid and a base, and are, therefore, salts. In all fats and oils the base is **glycerine**, but the acid varies. Thus in tallow the acid is stearic acid, in palm oil it is palmitic acid, and in olive oil it is oleic acid. Fats are solid, and oils are liquid, and both are insoluble in water. Nearly all the fats are derived from animals, and the oils are obtained chiefly from the seeds and fruits of vegetables.

247. Tallow.

Tallow is the name applied to the fat obtained from the ox, sheep, etc. The fatty substance of the animal, as it is removed from the body, is composed of minute fat cells, connected together in a mass by a membranous substance. The fat in the little cells is liquid during life, but solidifies after the animal is dead. By boiling the fatty substance in water till the fat is melted, and then putting it under pressure, the pure

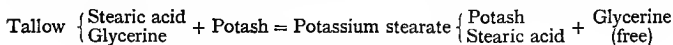
fat may be separated from the membranous matter which enclosed it. This fat is the tallow of commerce. As before stated it is a salt composed of the base glycerine with stearic acid, and may be called *glycerine stearate*. Tallow is lighter than water. When placed in hot water it melts, and floats on the surface.

248. Olive oil.

This oil is prepared from the fruit of the olive tree. It is lighter than water, and will not dissolve in that liquid. Its chief constituent is glycerine oleate, or olein—a compound of glycerine and oleic acid. It is extensively used in the arts and manufactures, and also for domestic purposes—for salads, for preserving fish, and as a medicine.

249. Action of alkalies on fats and oils.

Experiment 145.—Add a little tallow to hot water, and when it has all melted, add a solution of caustic potash. The tallow now disappears, and the whole of the liquid becomes slightly milky. The glycerine of the tallow has been displaced by the potash, and remains dissolved in the water as free glycerine. The potash, by combination with the stearic acid, has formed potassium stearate.



250. Stearic acid.

By adding hydrochloric to the solution of potassium stearate (soft soap) thus formed, the potash is again separated from the stearic acid; and the latter compound, being insoluble in water, separates out in the solid state. Crystals of stearic acid may be produced by dissolving this in alcohol or ether and evaporating. Stearic acid is extensively used in the manufacture of the so-called stearin candles, which do not melt at so low a temperature as those made from tallow.

251. Stearic acid and glycerine from tallow.

Tallow may be decomposed into stearic acid and glycerine by distilling at a temperature of 300° C. with steam under high

pressure. The stearic acid and the glycerine both pass over in the form of vapour with the steam, and when cold the bodies remain separate.

252. Summary.

1. Fats and oils are salts, insoluble in water.
2. All contain glycerine in combination with an acid.
3. Tallow is glycerine stearate.
4. Oils and fats are decomposed by alkalies, the glycerine being set free.
5. Tallow, heated with a solution of potash, forms potassium stearate and free glycerine.
6. Potassium stearate is decomposed by hydrochloric acid, the stearin being set free.
7. Stearin is soluble in alcohol and ether.

QUESTIONS ON LESSON XXX.

1. What are the chief properties common to fats and oils?
2. What do you know of the chemical composition of fats and oils?
3. What is tallow? How is it prepared?
4. Describe the chemical composition of tallow.
5. Describe fully the action of potash on tallow when heated with water.
6. What is stearin? How may it be prepared?
7. How may tallow be separated into glycerine and free stearic acid?

LESSON XXXI.

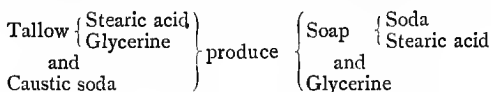
SOAP AND GLYCERINE.

253. How to make soap.

THERE are several varieties of soap, but they are all formed by the combination of an alkali with a fat or oil. Almost all the fats and oils have been used for the preparation of soap, and the alkali used is potash or soda. The following experiment illustrates the method of preparing the ordinary white or yellow soap.

Experiment 146.—Put some tallow in a beaker glass with a few ounces of water. Support the beaker on a piece of wire

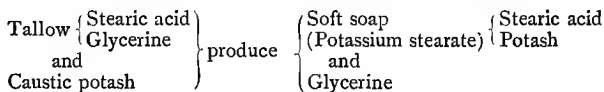
gauze, and apply heat. When the tallow has melted add a little caustic soda, and continue the heating very gradually to prevent the liquid from frothing over. After boiling gently for a short time remove the source of heat, and add a little common salt. The stearic acid of the tallow has now combined with the soda, and thus formed **sodium stearate**. The addition of common salt causes this sodium stearate, which is soap, to separate from the liquid, and as the mass cools, the soap solidifies on the surface. The change may be represented thus :—



The glycerine which is set free, remains mixed with the water, and may be separated and purified by an after process. The presence of stearic acid in soap may be proved by dissolving the soap in water, and adding hydrochloric acid. Sodium chloride is formed, and the stearic acid is thrown down as a white powder.

254. Soft soap.

Soft soap is composed of stearic acid and potash, and may be prepared by a process exactly similar to that illustrated by the last experiment, substituting caustic potash for the caustic soda. The soap thus produced is more powerful than the ordinary soap, and does not become hard on cooling.



255. Action of soap in washing.

All oils and fats are insoluble in water, therefore water alone will not cleanse greasy surfaces. But by the use of soap the fatty matter is dissolved and consequently is easily removed. Even in the washing of the skin we find that water alone is not very effectual in removing the impurities. Small quantities of fatty matter are continually exuding through the pores of the skin, and become mixed with any impurities which may be

present on the surface of the body. Hence it is that we find the use of soap so necessary for the thorough cleansing of the skin.

256. The action of soap on hard water.

Those who have used both hard and soft waters for washing must have noticed the comparative ease with which a lather may be produced by means of the latter with a little soap. In washing, it is necessary that the soap should be in solution. This solution is very readily obtained with soft water. But hard water contains the salts of lime ; and when soap is mixed with hard water, instead of forming a simple solution, it combines with the lime and forms an insoluble substance called **calcium stearate**, which floats on the surface in white masses. After all the lime present has thus been converted into the stearate, the soap begins to dissolve, and then its cleansing properties become apparent. It will thus be seen that hard water may be softened by using sufficient soap to convert all the lime into an insoluble form ; but this is an expensive and troublesome method, compared with the methods mentioned in Lesson II.

257. Glycerine.

Glycerine is a thick, syrupy, colourless liquid, having an intensely sweet taste. Its specific gravity is 1.2. It is composed of the same elements as sugar, but in different proportions. Glycerine is soluble in water and alcohol. It is decomposed when heated alone, and leaves a black mass of carbon ; but when heated with water it is unchanged ; the glycerine vapour passes over with the steam, and both may be condensed together in a receiver. If pure glycerine be reduced to a low temperature it becomes solid.

258. Preparation of glycerine.

Two methods of separating glycerine from fats have already been described (Arts. 253 and 254) ; but very large quantities of glycerine are manufactured by decomposing fat with lead oxide. The stearic acid combines with the lead oxide, and forms 'lead soap,' and the glycerine is thus set free.

259. Summary.

1. All soaps are compounds produced by the combination of fats or oils with alkalies.

2. The ordinary yellow soap is prepared by boiling tallow with caustic soda. It is sodium stearate.

3. Soft soap is potassium stearate. It is prepared by boiling tallow with caustic potash.

4. Oil and grease may be removed by washing with water containing soap, as this substance has the power of dissolving fatty matter.

5. Soap forms the insoluble calcium stearate when mixed with hard water.

6. When all the lime has been thus converted into the stearate the water is rendered soft.

7. Glycerine is composed of carbon, hydrogen, and oxygen.

8. It is prepared by

(a) Boiling a fat or oil with an alkali.

(b) Distilling a fat with high-pressure steam.

(c) Decomposing fat with lead oxide.

QUESTIONS ON LESSON XXXI.

1. What is soap? Describe its composition.

2. How is soap prepared from tallow?

3. What is soft soap? How does it differ in composition from the ordinary hard soap?

4. How is soft soap prepared?

5. Explain the use of soap in washing.

6. Describe fully the action of soap on hard water. How is it that a lather may be more easily produced by using soft water?

7. Describe the composition of glycerine.

8. How would you prove that glycerine contains carbon and hydrogen?

9. Give the chief physical properties of glycerine.

10. How is glycerine made?

LESSON XXXII.

SUGAR.

260. The chief sources of sugar.

THE sugar we use is derived from the juices of various plants. Sometimes it is abundant in the stem, sometimes in the root, and a variety of sugar is also met with in most fruits. Our supplies are obtained chiefly from the sugar-cane, the beet-root, and the sugar-maple. The solid part of honey consists of a kind of sugar.

261. Composition of cane-sugar.

Cane-sugar is composed of the elements carbon, hydrogen, and oxygen. It consists of 42.1 parts of carbon and 6.4 parts of hydrogen to 51.5 parts of oxygen.

262. Properties of cane-sugar.

Sugar melts at 160° C. If still further heated it blackens. This is due to the setting free of some of its carbon. If the heat be continued the sugar is entirely decomposed, and a black, porous mass of carbon remains. When pure, sugar is a white, crystalline compound, with a sweet taste, and is readily soluble in hot or cold water. One part by weight of cold water will dissolve three parts of sugar, but hot water will dissolve sugar in much larger proportions. The solution is neutral. When a hot and saturated solution of sugar is allowed to cool, it forms crystalline masses called *sugar-candy*. If heated to 180° C. with a little water, sugar forms on cooling a transparent mass of non-crystalline *barley-sugar*. Sugar will combine with the oxides of some metals, but not with acids. If, however, it is heated for some time with dilute acid, it is converted into other kinds of sugar called grape sugar and fruit sugar.

263. Grape sugar.

That kind of sugar which is found so extensively in fruits is called grape sugar, because it is present in large quantities

in the juice of the grape. This sugar contains the same elements that compose cane sugar, but in different proportions. In 100 parts there are 40 of carbon, 6.7 of hydrogen, and 53.3 of oxygen.

Experiment 147.—Put some honey into a glass and add about the same volume of alcohol. Stir well, and, after allowing the solid part to settle down, decant off the clear liquid. Let this be repeated with the residue and a fresh portion of alcohol. The sediment remaining is chiefly grape sugar, which, being only slightly soluble in water, may be finally washed with cold water. Grape sugar is not so sweet or soluble as cane-sugar.

There are other kinds of sugar, among which may be mentioned *milk sugar* and *fruit sugar*. The former is found in the milk of mammals, and may be obtained by evaporation; the latter is found in many fruits mixed with grape sugar.

264. Summary.

1. Sugars are derived from the juices of plants. All are composed of carbon, hydrogen, and oxygen.
2. Sugar is decomposed by heat, a mass of carbon remaining.
3. Sugar candy is sugar crystals deposited from a hot solution as it cools.
4. Cane-sugar is converted into fruit sugar and grape sugar by long boiling with dilute acid.
5. Grape sugar is not so sweet and not so soluble as cane-sugar.
6. The solid part of honey consists chiefly of grape sugar.

QUESTIONS ON LESSON XXXII.

1. Name some of the different kinds of sugar. What do you know of their composition?
2. Give the chemical composition of cane-sugar.
3. How would you prepare a crystalline mass of sugar?
4. Describe the chief properties of cane-sugar.
5. How would you prove that sugar contains carbon?
6. What is the action of acids on cane-sugar?
7. What is grape sugar? Describe its composition.
8. How would you prepare a small quantity of grape sugar?
9. How does grape sugar differ from cane-sugar?

LESSON XXXIII.

STARCH, DEXTRIN, AND GLUTEN.

265. The properties of starch.

STARCH is a vegetable compound found in all parts of plants, but especially in the roots and leaves. It is the chief constituent of potatoes, rice, sago, tapioca, and arrowroot. Starch is a neutral substance. It is not crystalline in structure, but consists of very minute granules or cells, which vary in size in different plants. Starch is insoluble in cold water, but when boiled with water the little cells swell up and burst, and some of the starch dissolves. When this solution cools it forms a thick paste, which is used for stiffening linen. If we boil some of the starch paste for some time with more water the starch is changed into a soluble form, and the solution becomes clear.



FIG. 63.—STARCH GRAINS, AS SEEN UNDER THE MICROSCOPE.

266. How to obtain starch from potatoes.

Experiment 148.—Cut a large potato in two, and rub the cut surfaces together in a fine stream of water issuing from a tap. Allow the washings to pass into a clean vessel and settle, then pour off the clear water. The sediment now remaining consists of starch cells, each single cell having a diameter of about $\frac{1}{160}$ th of an inch.

267. The composition of starch.

Starch contains the same elements which are found in sugar, in the following proportions by weight :—

Carbon	44.4
Hydrogen	6.2
Oxygen	49.4
	<hr/>
	100.0

268. Test for starch.

Starch, both in its soluble and insoluble forms, produces a blue colour with iodine. This colour may be destroyed by heating, but returns again on cooling. The presence of starch may be proved by the application of this test.

269. Dextrin.

Experiment 149.—Heat some starch to a temperature of about 150°C. , and it will be converted into a soluble substance known as dextrin, or British gum. The conversion will be much more rapid if a small quantity of dilute nitric or hydrochloric acid be added. Dextrin is soluble in hot and cold water, and is largely used as a substitute for gum arabic. Extract of malt will also convert starch into a soluble substance, which is really a mixture of dextrin and grape sugar.

270. Gluten.

This is a sticky, elastic substance, found in wheat flour. When fresh and moist it may be drawn out into threads. Unlike starch, it soon putrifies when exposed to the air, giving off a very unpleasant odour.

271. Composition of gluten.

Gluten contains the elements carbon, hydrogen, oxygen, and *nitrogen*, but it varies in chemical composition, as it consists of these four elements united together in slightly varying proportions.

272. How to obtain gluten.

Experiment 150.—Tie some wheaten flour in a calico bag, and well knead it for some time in a vessel of water. The water becomes milky through the separation of starch which passes through the calico from the flour. (Prove this to be starch by means of the iodine test.) Now examine the contents of the bag. These will be found to consist of a sticky mass, composed largely of gluten. By kneading for a very long time, all the starch may be removed. Wheat flour is composed of about 70 per cent. of starch and about 10 per cent. of gluten.

273. Food substances.

The animal and vegetable substances that we use for food may be divided into two classes :—*nitrogenous* and *non-nitrogenous*. The former class includes those foods which contain nitrogen, and the other class those foods which are composed of carbon, hydrogen, and oxygen only. The nitrogenous foods are sometimes called *flesh-forming foods*, as they contain the nitrogen which is an essential constituent of flesh or muscle. The non-nitrogenous foods are sometimes termed *heat-producing foods*, as they are supposed to be the chief source of heat in the body. Non-nitrogenous foods are also called *fat-producing foods*. Sugar, fats, and starch are examples of this class ; and gluten is one of the nitrogenous or flesh-forming foods.

274. Summary.

1. Starch is a vegetable compound, composed of carbon, hydrogen, and oxygen.
2. It is insoluble in cold water, but slightly soluble in hot water.
3. Starch is cellular or granular in structure.
4. It forms a blue colour with iodine.
5. When heated to 150° C. it is converted into dextrin, or British gum.
6. Gluten is found in wheat and other cereals. It contains carbon, hydrogen, oxygen, and nitrogen.
7. Foods may be classified as follows :—

Foods	{	Nitrogenous , containing carbon, hydrogen, oxygen, and nitrogen. Example—gluten.
		Non-nitrogenous , containing carbon, hydrogen, and oxygen. Examples—starch, sugar, fat.

QUESTIONS ON LESSON XXXIII.

1. Whence do we obtain starch ? Describe its chief properties.
2. How would you make a solution of starch ?
3. How would you obtain starch from a potato ?
4. Describe the structure of starch.
5. Give the chemical composition of starch. How would you prove that it contains carbon and hydrogen ?

6. By what means would you distinguish starch from other organic substances?
7. What changes take place when starch is heated?
8. What is dextrin? How may it be prepared?
9. Describe the properties and chemical composition of gluten.
10. How may gluten be prepared from wheat flour?
11. What do you know of the constituents of wheat flour?
12. What are the two classes of food substances? Give examples of each class.

LESSON XXXIV.

ALCOHOL, OR, SPIRIT.

275. The chief properties of alcohol.

ALCOHOL is a colourless, neutral, and volatile liquid, having a pleasant, spirituous odour, and a burning taste.

Experiment 151.—Dissolve a little aniline in alcohol, and pour the solution thus obtained very gently on water. The alcohol, being lighter than water (specific gravity = $\cdot 8$), floats on the surface of the latter, but dissolves slightly where the two liquids come in contact. If we now *mix* the liquids by stirring or shaking, the alcohol *dissolves* in the water, at the same time giving off heat.

Experiment 152.—Pour some alcohol into a test-tube and dip the tube into a vessel of water heated to near its boiling point. In a short time the alcohol will boil, giving off a colourless vapour which may be ignited as it rises. The boiling point of alcohol is 78° C.

The freezing point of alcohol is not known, as it has never yet been solidified. On this account, and also because it expands and contracts regularly, it is largely used in the manufacture of thermometers, especially those intended for the registering of extremely low temperatures.

Spirits of wine is alcohol which has not been perfectly freed from the water which it always contains when first prepared. By an after process, which consists of a second distillation with a substance, such as lime, having a stronger affinity for water than the alcohol has, the water is removed,

and **absolute alcohol** is obtained. As the duty on pure alcohol is very high, the Government allow a mixture of 90 parts of strong spirit and 10 parts of *wood-spirit* or *methyl-alcohol* to be sold free from duty for chemical and manufacturing purposes. This mixture is known as **methyiated spirit**.

276. The combustion of alcohol.

Experiment 153.—Pour a little alcohol into a clean watch-glass or evaporating dish and ignite it. It burns with a very pale blue flame. Hold over the flame an inverted bell jar. The jar is instantly covered with a deposit of dew, thus proving that water is a product of the combustion, and, therefore, that alcohol contains hydrogen. Now close the jar quickly by means of a glass plate, and shake up the air which it contains with a little lime water. The formation of chalk proves the presence of carbonic acid gas. Therefore the alcohol contains carbon. The composition of alcohol by weight is carbon 52·2 parts, hydrogen 13 parts, and oxygen 34·8 parts. When all the alcohol has burnt away, it will be noticed that no residue of carbon remains, but that it leaves the dish perfectly clean.

277. Alcohol as a solvent and preservative.

Alcohol is extremely useful as a solvent. Many substances used in the preparation of medicines are insoluble in water, but soluble in alcohol. Nearly all the varnishes and polishes used by mechanics are prepared by dissolving various gums and resins in spirit. The spirit evaporates rapidly, leaving a shining film of the solid substance which will not wash off with water. Alcohol is also a valuable preservative. Animal substances may be preserved indefinitely by covering them with spirit.

278. Intoxicating liquors.

Alcohol is the intoxicating principle of beers, wines, and spirits. Ale contains from 3 to 8 per cent. of alcohol; wines from 8 to 16 per cent.; and brandy from 40 to 60 per cent. All these consist almost entirely of spirit and water, the different flavours being due to the small quantities of other substances present. When a mixture of alcohol and water is

distilled, the alcohol passes off at first more readily than the water, and towards the end of the process the proportion of water is much larger. By mixing with lime before commencing the distillation, the water is kept back, and the alcohol may be made to pass off almost free from water.

Experiment 154.—Put a little beer into a glass retort,

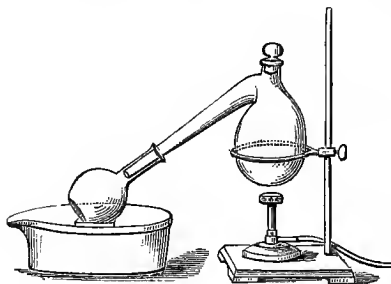


FIG. 64.—APPARATUS FOR EXPERIMENT 154.

apply heat, and condense the vapour given off in a flask kept cool by a running stream of water. Distil only a small quantity and throw away the remainder. Mix this distillate with quicklime, allow it to stand for an hour or two, and then distil again. The second dis-

tilate thus produced contains a large proportion of alcohol, has a burning taste, and may be readily ignited.

279. 'Proof spirit.'

The proof spirit of the Excise contains 50·8 parts of water and 49·2 parts of alcohol. This mixture, when ignited on gunpowder, will fire the latter. If a larger proportion of water be present it will not ignite gunpowder, and the mixture is then said to be '*under-proof*'.

280. Fermentation.

Fermentation is a peculiar chemical action due always to the presence of minute living beings, belonging generally to the vegetable kingdom. There are different kinds of fermentation, but one of the most common is that in which alcohol and carbonic acid are formed. Acetic and other acids are also produced by the action of certain ferments—ferment being the name given to the organism which has the power of producing the action. Sometimes, as in the case of beer and wine, the fermentation sets in without the addition of a ferment. But in

such cases either the liquid already contains a ferment, or the seeds or *sporules* of the living beings float about in the air and drop into it. During the fermentation of a liquid the rapid multiplication of the organisms may be watched under the microscope. The yeast plant is the ferment used in the brewing of beer. The juice of the grape already contains a ferment, so that no such addition is necessary in the making of wines. The following experiment illustrates a simple example of fermentation :—

Experiment 155.—Place a solution of sugar in a large glass flask and add some yeast. Let the flask be provided with a cork, and a delivery tube which dips into a solution of clear lime-water. Set this apparatus aside in a warm place for a few days. At the end of this time a deposit of chalk will be found in the vessel which contained the lime-water ; and alcohol may be obtained by the distillation of the liquid contained in the flask. The cane sugar was first changed into grape sugar by the action of the yeast, and then into alcohol and carbonic acid gas. Almost all the alcohol used in the manufacturing industries is prepared by the fermentation of sugar.

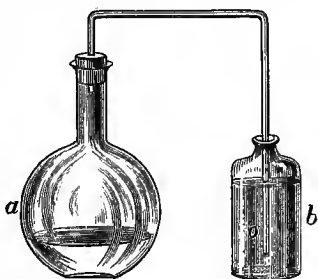


FIG. 65.—ILLUSTRATING EXPERIMENT 155.

a, Flask containing sugar, water, and yeast ; *b*, bottle containing lime-water.

281. Summary.

1. Alcohol, or spirits of wine, is a colourless, neutral, volatile, and combustible liquid. It is lighter than water, in which it dissolves, giving off heat.
2. Alcohol boils at 78° C., and does not solidify at the lowest known temperatures.
3. Methylated spirit is a mixture of common alcohol and methyl alcohol or wood-spirit.
4. Alcohol contains carbon, hydrogen, and oxygen. When it burns, water-vapour and carbonic acid gas are produced, and no residue is left.

5. Alcohol is a valuable solvent and preservative.
6. Alcohol is the intoxicating principle of beers, wines, and spirits. It may be separated from these liquids by distillation with lime.
7. Proof spirit contains 50.8 parts of water and 49.2 parts of alcohol.
8. Fermentation is a chemical action, initiated and carried on by a substance called a ferment.
9. The presence and growth of certain living beings determine this action.
10. Alcohol and carbonic acid gas are produced by the fermentation of sugar.

QUESTIONS ON LESSON XXXIV.

1. Describe a few simple experiments which illustrate the chief properties of alcohol.
2. How would you prove that the boiling point of alcohol is lower than that of water?
3. What are the freezing and boiling points of alcohol?
4. What are *spirits of wine* and *methylated spirit*?
5. What is *absolute alcohol*? How is it obtained?
6. How would you prove that alcohol contains carbon and hydrogen?
7. Name some substances that are soluble in alcohol, but insoluble in water.
8. Name some substances that are soluble in both water and alcohol.
9. What are the chief uses of alcohol?
10. What do you know of the composition of beer, wine, and spirits?
11. How would you obtain pure alcohol from wine?
12. What is '*proof spirit*'? By what test may we know whether spirit is '*over proof*' or '*under proof*'?
13. What is *fermentation*? What changes take place during fermentation? What conditions are necessary for the production of these changes?
14. Describe an experiment which illustrates the action called fermentation.
15. How is alcohol prepared?

EXAMINATION QUESTIONS

*SET BY THE DEPARTMENT OF SCIENCE AND ART
AT THE MAY EXAMINATIONS.*

1887.

1. A glass of water is exposed to the air ; in time the water disappears into the air : how do you account for this ? How could you prove that there is moisture in air ? (13.)
2. Air is passed over red-hot iron : what change does this cause in the air and in the iron ? (9.)
3. How could you show that the gas obtained by dissolving marble in hydrochloric acid is also contained in the breath ? (10.)
4. Two samples of water are given to you ; one is a hard water, and the other is distilled water : describe two methods of distinguishing between them. (13.)
5. What is vinegar ? How is it prepared ? Vinegar is poured upon washing soda : what happens ? (9.)
6. Ammonia is classed as an alkali : why ? Name some of the sources from which it can be obtained, and give its composition. (15.)
7. A piece of lead, a piece of copper, and some mercury are separately heated in a crucible over a lamp : describe what occurs in each case. (11.)
8. From what substances can starch be obtained ? Of what is it composed, and how does it behave when boiled with water ? (13.)
9. Name some commonly occurring compounds of sodium. How can you show that chlorine is a constituent of common salt ? (10.)
10. What substances are contained in flour ? How can they be separated, and what essential difference is there in their composition ? (12.)
11. What is meant by saying that a solution is saturated ? How would you prove that no loss of weight occurs when a substance is dissolved in water ? (8.)
12. What are the distinguishing characters of cast iron, wrought iron, and steel ? What is iron rust ? (13.)

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